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To President Northrop,
with regards,
H. C. Carel.

Inorganic Chemistry Syllabus

BY

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INTRODUCTORY DEFINITIONS.

Science—

Classified knowledge and deduced relation; natural science objective not subjective.

Divided—

Into two classes which shade into each other:

Animate	{ Biology	{ Botany	Inanimate	{ Chemistry
		{ Zoology		{ Physics
				{ Mineralogy, etc.

Physics and Chemistry—

Treat of changes of matter; *i. e.*, that which occupies space.

Physical—

Change in place, condition or properties, substance unchanged.

Chemical—

Change of substance and properties.

Chemistry—

Treats of what substances are composed, what occurs when they change and the dependence of properties on composition—chemical knowledge depends on analysis, and synthesis.

Analysis—

Chemical decomposition of complex, to simple substances.

Synthesis—

Chemical combination of simple to complex substances.

Element—

A substance as yet undecomposed. All matter is formed from one element or a combination of elements.

Allotropic—

Modification of an element is a substance of same elementary composition but differing in chemical and physical properties.

Isomerism and Isomers—

The above phenomenon in compounds is termed isomerism and such bodies are isomers.

Physical—

Form of elements—gases, liquids and solids.

Chemical—

Division—metals and non-metals.

Metal (*in general*)—

Chemical element which unites with the elements of water to form a base (p. 15). Except hydrogen, metals in general are opaque solids, more or less malleable, ductile and tenacious, conductors of heat and electricity and possessed of a peculiar lustre termed metallic.

Non-Metals (*in general*)—

Are, as the name suggests, the direct opposites of metals, *i. e.*, they unite with elements of water to form an acid. Three of the non-metals are gases: Oxygen, Nitrogen, and Chlorine—one probably a gas, Fluorine—one is a liquid, Bromine—the rest are solids.

Compound—

Union of two or more elements in simple proportions, to form a new substance.

Mixture—

Mechanical intermingling of matter in any or all proportions, without change of substance. The original materials may be recovered by mechanical means.

Chemical Affinity—

Tendency of elements to unite—C. A. of a given element may vary in uniting with different elements, but is always constant with any particular element. Some elements unite directly to form a compound $A+B=AB$. More often an element unites with a compound $A+CD=AD+C$; or two compounds react to form two new compounds $AB+CD=AC+BD$.

Metathesis—

The last reaction is known as metathesis (*Gr.*, to set over), and occurs especially in solution, where one of the products is insoluble, or by heat, where one of the products is volatile.

Heat and Chemical Energy—

Save nitrogen, all union of elements yields heat; that is,

chemical energy transforms to heat. Decomposition of compounds absorbs heat which is transformed into chemical energy.

Acid (*in general*)—

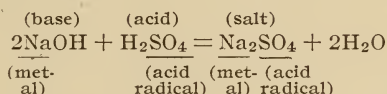
Compound of hydrogen which yields salts by replacing its hydrogen with a metal. Acid reddens blue litmus and possesses characteristic “acid” taste. That part of acid which unites with the metal is called “acid radical.”

Base (*in general*)—

Oxide or hydroxide (OH) of a metal which exchanges its metal for the hydrogen of an acid forming a salt. Base blues litmus and possesses characteristic “alkali” taste.

Salt—

The generally neutral result of the combination of metal of base with radical of acid.



Neutralization—

This reaction of an acid body with alkali forming a substance neither acid nor alkaline is termed neutralization.

LAWS OF CHEMICAL PROPORTIONS.

Composition of a given substance is invariable.

Definite Proportions—

Chemical combination always takes place between definite masses. Any excess is unacted on.

Multiple Proportions—

When elements unite in more than one proportion, the ratio of succeeding compounds are simple multiples of the first ratio.

Example: $\text{N}_2\text{O}—\text{N}_2\text{O}_2—\text{N}_2\text{O}_3—\text{N}_2\text{O}_4—\text{N}_2\text{O}_5$.

Atomic Theory.

Above laws depend on the Atomic Theory that every body is an aggregate of atoms or ultimate indivisible particles. The absolute weight of an atom is of course unobtainable, but we have relative atomic weights with hydrogen as the unit.

Atom vs. Molecule.

Atom = smallest indivisible particle.

Molecule = smallest divisible particle.

Law of Charles—

Volume of gases varies directly as the absolute temperature.

Law of Mariotte—

Volume of gases varies inversely as pressure.

Law of Avagadro—

At same temperature and pressure equal volumes of all gases contain the same number of molecules.

Certain Physical Properties—

Of bodies are always the same for the same substance, *i. e.*:

- (1) The boiling point of liquids.
- (2) The melting point of solids.
- (3) The crystalization of solids.

SOLUTION.

Some substances require particular solvents, as ether for fats, carbon disulphide for yellow phosphorus, etc., but the great solvent is water. In studying the phenomena of solution we notice, (1) the uniform distribution of dissolved substance, (2) the almost infinite subdivision of material, (3) that substances vary as to solubility, and that the presence of foreign matter often affects solution.

(4) Evaporate a solution and (a) solids remain, (b) gases escape, (c) liquids pass off at their boiling points.

(5) Solution is of great advantage in chemical reactions where often the new substance is insoluble and can be separated.

**GROUPING OF ELEMENTS WITH REGARD TO
PROPERTIES.**

Non-Metals—

Group I—Cl, Br, I, F.

Group II—O, S, Se, Te.

Group III—N, P, As, Sb, Bi.

Group IV—B.

Group V—Si, C.

Metals—

Group I—H.

Group II—K, Rb, Cs, Na, Li.

Group III—Ca, Sr, Ba.

Group IV—Be, Mg, Zn, Cd, Hg.

Group V—Pb, Tl, Cu, Ag, Au.

Group VI—Y, La, Cr, Di, Sr.

Group VII—Al, Ga, In.

Group VIII—Mn, Fe, Co, Ni, Cr, Mo, W, U.

Group IX—Sn, Ti, Zr, Th.

Group X—V, Ta, Nb.

Group XI—Pt, Ru, Rh, Pd, Ir, Os.

OXYGEN (to generate acid).

At. Wt. 16. Val. II. Symbol O.

Occurrence—

$\frac{8}{9}$ of water, $\frac{1}{5}$ of air, $\frac{1}{2}$ earth's crust, $\frac{3}{4}$ organized matter.

Preparation—

By heating (1) mercuric oxide HgO , (2) manganese dioxide MnO_2 , (3) potassium chlorate KClO_3 . (4) Decomposition of water (H_2O) by electricity. (5) In the laboratory, by heating a mixture of $\text{KClO}_3 + \text{MnO}_2$. (6) Commercially obtained by heating barium oxide BaO to a red heat in the air. This forms barium peroxide BaO_2 , which heated to a white heat loses half its oxygen and is reduced again to the oxide.



Properties—

Colorless, inodorous, tasteless gas—the great supporter of combustion and life—combines with all elements save fluorine to form oxides. At ordinary temperatures oxygen is moderately active, but its chemical affinity increases with heat. O is slightly soluble in water—with low temperature and great pressure condenses to a liquid.

Detection—

Distinguished from all gases but nitrous oxide N_2O by its spark. From N_2O (1) by the solubility of O in potassium pyrogallate, (2) Sulphur burns in O but not in N_2O , (3) O, exploded with two volumes of hydrogen (H), loses all its volume, while N_2O leaves a residue of N.

Combustion—

Rapid oxidation accompanied by light and heat.

Decay—

Slow oxidation of organic matter assisted by bacteria.

Kindling Point—

Temperature at which bodies combine with light and heat
—Interposition of such bodies as wire gauze serves to reduce the temperature below the kindling point. In DAVY SAFETY LAMP, a wire gauze jacket surrounds the flame. The heated gases passing through such a jacket are cooled below the kindling point, and there can be no combustion. Small quantities of “fire damp” passing through the gauze from without, are heated within the jacket to their K. P. and cause a small explosion which warns of a dangerous proximity.

Flame—

A burning gas whose luminosity is usually due to red hot un-oxidized carbon (C). Intense light is produced by the glowing of some non-volatile bodies in a flame as calcium oxide (CaO) in calcium light. Other things being equal, combustion produces less light and more heat in proportion as oxygen is increased. In illuminating tips, gas burns by oxygen obtained at tip of tube, while in the colorless “Bunsen” the same gas is well mixed with air before it ignites. In the oxyhydrogen lamp or blow-pipe Oxygen alone is mixed with Hydrogen giving an intense non-luminous heat. The comparative heat of burning bodies is known and Carbon (C) yields the best fuel by volume.

Oxidation—

Adding of O to any substance.

Reduction—

Subtracting of O from any substance.

Reactions of Oxides—

Oxides may have one of three reactions to litmus.

Acid—	Basic—	Neutral—
SO ₂ ,	K ₂ O,	H ₂ O,
NO ₂ ,	Na ₂ O,	MgO,
etc.	etc.	etc.
Red.	Blue.	Purple.

OZONE, O₃.

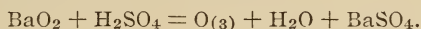
Oxygen occurs in two allotropic forms, ordinary oxygen O₂ and ozone O₃.

Occurrence—

Whenever oxygen is prepared, small amounts of ozone form. It is supposed to exist in the air, but we have no proof.

Preparation—

(1) By the silent discharge of an electric machine. (2) Slow oxidation of phosphorus. (3) In the laboratory by treating barium peroxide with sulphuric acid.



Properties —

“Condensed oxygen” is a heavy gas—poisonous—taste and odor like weak chlorine—at ordinary temperature is a powerful oxidizing, bleaching and disinfecting agent. Ozone acts with its extra atom of O—O₃ = O₂ + O. Here O does the work, and O₂ (ordinary oxygen) is set free. Ozone condenses to a blue liquid—at 300° breaks to O₂.

Detection—

(1) O₃ blues potassium iodide (KI) starch paper (common to Cl, NO₂ and H₂O₂). (2) Blues red KI litmus paper. This is distinctive in the absence of ammonia (NH₃). Chief use of O₃ is to bleach old pictures.

HYDROGEN—H.

At. Wt. 1. Val. I.

History—

Discovered by Cavendish in 1766.

Occurrence—

Free in volcanoes, oil-wells, and from decomposing organic matter; chiefly combined with oxygen in H₂O and organic bodies.

Preparation—

(1) From H₂O by (a) electricity, (b) metallic sodium or

potassium, (c) hot iron. (2) Laboratory method, by replacing H of an acid by a metal.



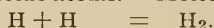
Properties—

Tasteless, inodorous, colorless gas—does not support combustion, but burns with a blue flame to H_2O —mixed with oxygen explodes at K. P.—lightest of gases, H is the unit of specific gravity for gases—highly diffusible and a strong reducing agent—H is a gaseous metal playing the same part in acids as the other metals play in salts—conducts heat and electricity, forms alloys, and in electric decompositions, goes to the negative (metal) pole.

Nascent gas—

Condition at immediate generation when chemical affinity is much stronger than usual—the atoms have not yet united to form the molecule.

Nascent atoms. — Molecule.



WATER— H_2O .

At ordinary temperature a transparent fluid devoid of taste or smell—thin layers colorless, large masses blue—cooled to a certain temperature water crystallizes, forming ice—this is the 0°C .—heated to a certain temperature water boils—this is the 100°C . On cooling water contracts till 4°C ., which is its maximum density—below 4° it expands, hence ice floats, and water pipes break. H_2O evaporates at all temperatures, ice vaporizes slowly and without changing to liquid form. Heated water changes rapidly to an invisible gas called STEAM, which partially recondenses to the liquid state when in contact with air.

Composition of water—

The electric current resolves water into two volumes of H and one of O; conversely, by the electric current—or by heat—two volumes of H and one volume of O always combine to H_2O without residue, and the H_2O thus formed may be heated to two volumes of steam. Thus water by volume = $\text{H}_2 + \text{O}$, or by weight = 2 H to 16 O, or 1 H to 8 O.

Distilled water—

Natural waters contain many substances in solution and to obtain water free of solids it must be boiled and the steam condensed, *i. e.*, distilled.

Water of Crystallization—

Molecules of water which enter into definite combination with many salts, when these crystallize from their water solutions. It often influences the color, and crystalline form, but may be driven off by heat without alteration of substance.

Efflorescence—

Some salts yield their water of crystallization to the air as $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, etc.

Deliquescence—

Some salts take up water from the air and ultimately melt, as CaCl_2 , NaNO_3 , etc.

Unit of density—

Water at 15°C . is unit of density for solids and liquids.

Hydrometer—

An instrument for determining specific gravity of liquids with reference to water (1000)

HYDROGEN PEROXIDE— H_2O_2 .

History—

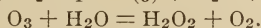
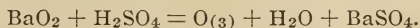
Discovered by Thenard in 1818.

Occurrence—

In small quantities in the air and formed wherever ozone oxidizes in presence of H_2O .

Preparation—

When barium peroxide BaO_2 is treated with H_2SO_4 a double reaction takes place. Ozone $\text{O}_{(3)}$ is generated, which then oxidizes the water formed to H_2O_2 .



Properties—

“Ozonic ether,” “golden fluid,” “a oxygenated water,” H_2O_2 is a thick colorless liquid. Soluble in H_2O —astringent

taste—dilute Cl odor—strong oxidizing and disinfecting agent, contains more O than any other substance (94%), concentrated solution easily breaks to $\text{H}_2\text{O} + \text{O}$ —dilute slightly acid solution is more stable.

Detection—

(1) Blues KI starch paper.

(2) When shaken with potassium bichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) and ether (C_2H_5)₂ O it gives the latter a blue color.

Use—

Antiseptic spray, and to bleach old pictures and hair.

CHLORINE—Cl.

At. Wt. 35.5. Val. I.

History—

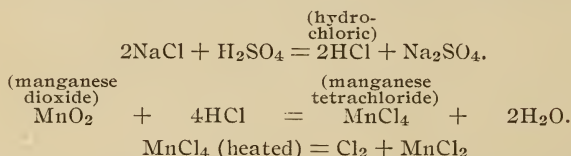
Discovered by Scheele in 1774.

Occurrence—

Chiefly as common salt NaCl in sea waters, springs, and especially rock salt deposits. Never free, because of strong affinities.

Preparation—

From NaCl, H_2SO_4 and MnO_2 a triple reaction.



breaking up of MnCl_4 shown by change of color—gas collected by displacement, or over hot water or salt solution.

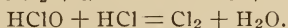
Properties—

Green yellow gas—violent odor, poisonous; strong chemical agent at ordinary temperature. For example, Cl extracts H from such hydrocarbons as turpentine—with some metals as antimony Sb it combines with light and heat; with H it unites in direct sunlight, forming HCl, or Cl will burn in H gas to form HCl and *vice versa*. Cl is a strong disinfecting agent and indirectly bleaches through the action of nascent O which it frees

from H_2O . Cl is easily soluble in H_2O , forming a solution which retains the properties of Cl gas. Cl used in commerce chiefly as bleaching powder.

Bleaching Powder (see Ca)—

Is formed by running Cl gas over slaked lime $\text{Ca}(\text{OH})_2$ —its approximate composition is $\text{CaCl}(\text{OCl})$. It disinfects and bleaches chiefly through nascent chlorine which is freed by dilute acids.



Detection—

By its odor, bleaching power, and by turning KI starch paper blue.

CHLORINE OXIDES.

Chlorine forms a series of oxides similar to that of nitrogen, *i. e.*, Cl_2O , ClO_2 , Cl_2O_3 , Cl_2O_5 . All are formed indirectly and unite with water to form acids.

CHLORIC ACID— HClO_3 .

Preparation—

The anhydride Cl_2O_5 is not yet known in the free state. The K salt of the acid is formed by passing Cl into KOH ,



The K salt is then treated with fluosilicic acid H_2SiF_6



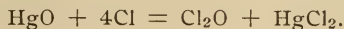
Properties—

A liquid of faint odor; strongly acid reaction; easily decomposed; a strong oxidizing and bleaching agent. The K salt is a source of oxygen supplying enough to burn some combustible substances. WHITE GUNPOWDER is composed of KClO_3 and sugar.

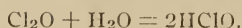
HYPOCHLOROUS ACID— HClO .

The anhydride Cl_2O is a red yellow gas with a Cl odor; explosive; condenses to a liquid; is prepared by the action of Cl on HgO .

mercuric oxide



The gas dissolves in water forming HClO



The Ca salt $\text{Ca}(\text{ClO})_2$ is highly important being the active principle of "bleaching powder." Hypochlorites in general are very unstable and act as strong oxidizing agents. The other oxides of Cl are unimportant.

HYDROCHLORIC ACID— HCl .

Occurrence—

HCl found free, in the gastric juice and in volcanoes—combined, chiefly as NaCl .

Preparation—

Generally prepared by action of sulphuric acid on common salt.



May be formed by direct union of H and Cl in the sunlight or by heat or electricity.

Properties—

Absolute HCl is a colorless, suffocating gas—non-combustible, and non-supporter—fumes in damp air—is eagerly absorbed by water, forming a 33 per cent solution which is the concentrated HCl of the laboratory—with silver nitrate (AgNO_3) hydrochloric acid forms a curdy, white, insoluble precipitate of silver chloride, AgCl , which is the test for HCl .

BROMINE— Br .

At. Wt. 80. Val. I.

History—

Discovered by Balard in mother liquor of salt works, 1827.

Occurrence—

Never free—usually as K , Na or Mg bromide—in salt water (1 gr. per gal.), rock salt deposits, sea-weeds and springs—accompanies chlorine and is obtained from mother liquor of salt works— Br is one of the less common elements.

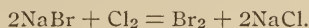
Properties—

Dark red liquid—red brown vapor—irritant poison—character similar to Cl but with weaker affinities being freed from

its compounds by Cl—Chemical unions of Br often accompanied by light—Br dissolves in water, alcohol, ether and chloroform.

Preparation—

Freed from its salts by (1) chlorine.



(2) Manganese dioxide and sulphuric acid.



(Same stages in this reaction as in preparation of Cl.)

Detection—

Add Cl water and CS_2 = red brown color to CS_2 . Br is set free by Cl— CS_2 dissolves free Br with a red brown color. Br gives starch an orange color.

Use—

Br is used in dye factories, photography, medicine.

HYDROBROMIC ACID—HBr.

Properties—

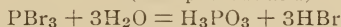
Colorless gas—fumes in moist air, eagerly absorbed by water, forming $\text{HBr} \cdot 2\text{H}_2\text{O}$. Analogous to HCl, but less stable, is decomposed by H_2SO_4 . Bromides freed by chlorine.

Preparation—

(1) Direct union of H and Br at red heat.

(2) Generally prepared by action of phosphorus tribromide and water.

(Phosphorous acid)



(3) Cannot be made pure from salt by H_2SO_4 , as resulting HBr is decomposed by H_2SO_4 .

Solubility—

Most Bromides are soluble in H_2O

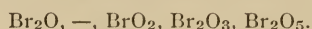
OXIDE AND CHLORIDE OF Br.

Bromine chloride, BrCl—

Br absorbs Cl forming a yellow unstable liquid with bleaching properties. Above 10°C . changes to Br and Cl.

Bromine oxides—

Bromine and oxygen unite with difficulty to form a series entirely analogous to Cl and made in same way—



IODINE—I.

At. Wt. 127—Val. I.

History—

Discovered by Courtois in 1811.

Occurrence—

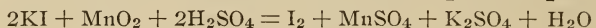
Not free—compounds accompanying Cl and Br, in sea-water, springs, etc.—Chief source is the ash of sea-weeds called “kelp.”

Properties—

Grey black solid with violet vapor—weak Cl odor—difficultly soluble in water—dissolves easily in alcohol or a solution of KI—chemically similar to Cl and Br but with weaker affinities, is freed by either from its compounds.

Preparation—

Kelp solution is partly evaporated—I remains in the mother liquor—solution is then distilled with MnO_2 and H_2SO_4 .



(Same reaction as for chlorine).

Detection—

- (1) With Cl water and CS_2 gives violet-color.
- (2) With cold starch paste, a blue color.

Used—

In dye factories and photography.

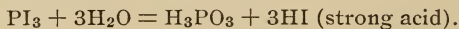
HYDRIODIC ACID—HI.

Properties—

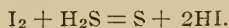
Colorless gas with suffocating odors—fumes in moist air—absorbed by water—easily decomposed; hence, cannot be made with H_2SO_4 —Strong reducing agent.

Preparation—

- (1) From Phos-tri-iodide and water.



- (2) I with H_2S gives weak acid.



OTHER COMPOUNDS OF IODINE.

Iodine Chlorides (ICl , ICl_3)—

Two compounds ICl , ICl_3 formed by direct action of the elements; used to add I to organic bodies.

Iodine Bromide (IBr)—

IBr is a solid similar to I and formed as ICl .

Iodine Oxides—

Iodine forms the same oxides as Br or Cl; I_2O , IO_2 , I_2O_3 , I_2O_5 .

Iodic Acid (HIO_3)—

A white solid formed by boiling I in HNO_3 ; permanent in air, but with heat changes to $\text{I}_2\text{O}_5 + \text{H}_2\text{O}$; forms salts similar to chlorates.

FLUORINE— Fl .

At. Wt. 19. Val. 1.

History—

Discovered by Davy in 1812, but not isolated till 1886 by Moissan.

Occurrence—

Chiefly in fluor spar or fluorite, CaF_2 , and cryolite AlFl_3 , 3NaFl .

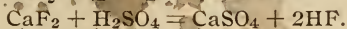
Properties—

Colorless gas—deadly poison—attacks most substances, but does not combine with O, C, Pb or Pt, nor attacks gutta percha—properties analogous to chlorine but much stronger, its atomic weight being much less. F frees Cl, Br or I from their compounds.

HYDROFLUORIC ACID, HF .

Preparation—

From fluor spar by H_2SO_4 in Pt retort.



Properties—

Colorless, mobile liquid—fumes in air—corrosive poison—dissolves all metals except Pb, Pt and Au, and all oxides, in-

cluding SiO_2 . Must be kept in gutta percha bottles. Used to etch glass and detected by its etching quality.

SULPHUR—S.

At. Wt. 32. Val. II.

History—

Used by the ancients for fumigation and medicine—Considered by alchemists to be the principle of combustion.

Occurrence—

- (1) Native in volcanic countries, especially Sicily.
- (2) Sulphates and Sulphides.
- (3) Organic compounds.

The amount of S produced is about 375,000 tons, of which nine-tenths comes from Sicily.

Extraction from Ores—

- (1) Melting ores in a pot, S floats and is dipped out.
- (2) Burning ores in small supply of air so that most S melts.

Purified—

By distillation—That which passes over first cools more rapidly forming FLOWERS and is less pure than the ROLL which forms afterwards—Flowers contain H_2SO_4 ; H_2S and SO_2 as impurities.

Properties—

Strong affinities—Union with metals often produces light and heat. Moist $\text{Fe} + \text{S} = \text{FeS} + \text{heat}$ —unites with haloids at ordinary temperature—Slightly heated $\text{S} + \text{P}$ explodes—H run into melted $\text{S} = \text{H}_2\text{S}$; Compounds of S alone are called sulphides. S is wholly analagous to O and when oxides of an element are soluble, usually sulphides are. With O, S burns with blue flame to SO_2 .

Uses—

Disinfectant; making H_2SO_4 , SO_2 , matches and gunpowder.

HYDROGEN AND SULPHUR.

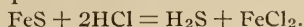
Two compounds well known—hydrogen sulphide, H_2S , and persulphide, H_2S_2 . The latter is unimportant.

Hydrogen Sulphide (H_2S)—Occurrence—

Free in mineral springs, and when organic bodies decay.

Preparation—

- (1) From a sulphide and an acid—



- (2) Chemically pure—



Collect over warm water or salt solution.

- (3) Formed by heating elements to $400^\circ C$.

Properties—

Colorless gas—bad odor—poisonous—combustible, burns to SO_2 and H_2O —mixture with air explodes—unstable, oxidizes in air at ordinary temperature to $S + H_2O$. Easily unites with metals to form sulphides. Dissolves in H_2O (3 volumes to 1), and water solution has properties of the gas, decomposing in light to $H_2O + S$. H_2S is a dibasic acid, hence forms two series of Salts, viz., KHS = potassium sulphhydrate, K_2S = potassium sulphide.

SULPHUR HALOIDS.

S unites with Cl in three proportions— S_2Cl_2 (SCl), SCl_2 , SCl_4 . The monochloride S_2Cl_2 is a red yellow liquid formed by passing Cl over molten S—used in vulcanizing rubber; SCl_2 and SCl_4 are highly unstable.

With Br and I, S forms similar compounds.

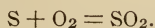
SULPHUR DIOXIDE, SO_2 .

Occurrence—

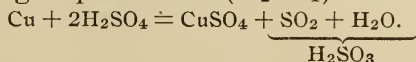
Native in volcanic gases.

Preparation—

- (1) Oxidizing sulphur—



- (2) Reducing sulphuric acid (H_2SO_4)—



SO_2 is called "Sulphurous acid" because the acid proper, H_2SO_3 , is unstable.

Properties—

Colorless, suffocating gas—inhibits combustion—poisonous to plants and animals—disinfecting agent—bleaches, but color is restored by alkalis—easily condensed without pressure—highly soluble in H_2O (50 volumes to 1)—salts easily oxidize to sulphates, hence good reducing agents. H_2SO_3 is a weak, dibasic acid, little stronger than carbonic— SO_2 does not unite directly with O but with ozone forms SO_3 .

SULPHUR TRIOXIDE— SO_3 .

Preparation—

Sulphuric anhydride formed from SO_2 by (1) ozone, (2) O over divided Pt, (3) electric spark, (4) distilling disulphuric acid, $\text{H}_2\text{S}_2\text{O}_7$.

Properties—

White, silky, crystalline needles—fumes in damp air—acid only with H_2O . Decomposes at high temperature to $\text{SO}_2 + \text{O}$. Unites with H_2O with great heat.

SULPHURIC ACID— H_2SO_4 .

History—

Known for centuries—alchemists prepared it from S and nitric acid (HNO_3) or by heating FeSO_4 , calling it “oil of vitriol”—first lead chambers built in 1746.

Occurrence—

Free in volcanic rivers and the fluids of some mollusca—combined as sulphates in large quantity as CaSO_4 , etc.

Preparation—

(1) Oxidation of S by HNO_3 directly.

(2) Oxidation of SO_2 by HNO_3 in presence of H_2O .— SO_2 from burning FeS_2 (iron sulphide), with O from the air, is passed into lead chambers, through which HNO_3 and H_2O are streaming— HNO_3 is decomposed to the oxides of N—these yield up O to SO_2 forming SO_3 , which unites with H_2O to form H_2SO_4 . The lower oxides of N reunite with O of the air, forming the higher, NO_2 and N_2O_3 , which again oxidize a new portion of SO_2 .

The dilute chamber acid (sp. gr. 1.5(c)) is concentrated in

Pt or glass vessels to commercial strength (1.830). Crude H_2SO_4 contains as impurities PbSO_4 , HNO_3 , often As from S or FeS_2 used. Impurities are generally removed by distillation, but an arsenic-free acid is obtained only by using arsenic-free materials.

Properties—

Heavy, syrup-like liquid, colorless when pure, usually brown from charred organic matter—odorless—corrosive—intensely acid. Heated above boiling, gradually breaks up to $\text{H}_2\text{O} + \text{SO}_3$ —mixed with water great heat is evolved, due to formation of higher hydrates— $\text{H}_2\text{SO}_4 + \text{H}_2\text{O}$ or $2\text{H}_2\text{O}$. Sulphates of strong bases are not easily broken up by heat.

FUMING OR NORDHAUSEN ACID— $\text{H}_2\text{S}_2\text{O}_7$.

By running SO_3 into concentrated H_2SO_4 we get $\text{H}_2\text{S}_2\text{O}_7$, disulphuric acid, a white solid stronger than H_2SO_4 and used in indigo and alizarin manufacture.

SELENIUM (Se) AND TELLURIUM (Te).

At. Wt. 79.

At. Wt. 128.

Two rare elements closely analogous to Sulphur—Tellurium was discovered in 1782 in ores of gold—physical properties are metallic, chemically it acts like S.—Selenium discovered in 1817 by Berzelius from the mud on floor of lead chambers. HNO_3 with Te or Se forms respectively tellurous acid, H_2TeO_3 or selenious acid, H_2SeO_3 —with KNO_3 the elements form K_2TeO_4 or K_2SeO_4 . Selenium colors the flame blue and H_2Se is remarkable for its odor. Selenium changes its resistance to an electric current when exposed to light.

SULPHUR GROUP.

O 16—S 32—Se 79.4—Te 128.

O a gas—others solids—O and S are non-metals—Se partly metallic—Te a metal. All unite with H_2 —the three solids form with H_2 volatile strong smelling gases of acid nature—Stability of H compounds, varies inversely as atomic weights, of O compounds directly—O is always bivalent, the others vary in valence

NITROGEN—N.

At. Wt. 14. Val. III. or V.

History—

Discovered in 1772 by Rutherford—Scheele and Lavoisier showed that air = O + N, called azote by the French and Italians.

Occurrence—

$\frac{4}{5}$ of the air and in many organic substances.

Preparation—

(1) Removing O from air by (a) Phosphorus—(b) red-hot Cu.

(2) Heating Ammonium nitrite $(\text{H}_4\text{N})\text{NO}_2$
 $(\text{H}_4\text{N})\text{NO}_2 = \text{N}_2 + 2\text{H}_2\text{O}.$

Properties—

Characterized by its inertness at ordinary temperature—colorless—inodorous, tasteless gas—non-combustible—non-supporter—not poisonous—at white-heat unites with metals—many of its compounds are explosives, while all explosives contain N.

AIR.

A comparatively constant mixture of $\frac{4}{5}$ N, $\frac{1}{5}$ O, with appreciable quantities of H_2O and CO_2 and trace of NH_3 compounds, NaCl, and dust. N and O are quite constant; the others vary.

Impurity of air is usually determined by amount of CO_2 present. CO_2 of itself is not harmful in small quantity, but experience has shown that it bears a constant ratio to the amount of ammonia compounds present. This “organic waste,” so-called, is intensely poisonous, although it exists in quantity too small for easy determination.

Air has the same physical properties as the gases of which it is composed, being a mixture, not a compound, because (a) it varies somewhat in composition, (b) proportion of O and N is not governed by atomic weight, (c) absorbed by water as free gases 35 vol. O : 65 vol. N, not 1 vol. O : 4 vol. N.

AMMONIA— NH_3 .

History—

Called “spirits of hartshorn” because formerly obtained

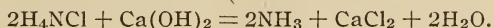
from horns of deer. Gas discovered by Priestley, who called it alkaline air.

Occurrence—

Not free in nature as NH_3 . Found as salt in animal secretions, especially urine. Found also when nitrogenous organic bodies decay or are distilled, hence found in liquor of gas works, as coal contains N compounds.

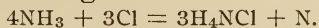
Preparation—

Commercially obtained from gas works. In laboratory from ammonium chloride (H_4N) Cl and calcium hydrate $\text{Ca}(\text{OH})_2$.



Properties—

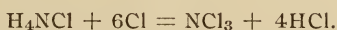
Colorless gas—pungent odor—strong alkali—non-supporter of combustion—burns feebly in the air. Decomposed by electricity or by chlorine in sunlight.



NH_3 is highly soluble in H_2O , forming H_4NOH , ammonium hydrate, a strong base in which the radicle, H_4N , departs itself like the metals Na or K, and will be treated with them.

NITROGEN HALOIDS

Do not form directly from elements—all are highly explosive—most important is NCl_3 —a dark red liquid, formed by action of Cl on H_4NCl .



NBr_3 and NI_3 are similar.

NITROGEN OXIDES.

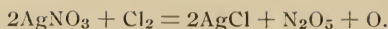
Five known compounds—

Nitrous oxide. Nitrous anhydride. Nitric oxide. Nitrogen peroxide. Nitric anhydride
 N_2O . N_2O_3 . $\text{NO}(\text{N}_2\text{O}_2)$ $\text{NO}_2(\text{N}_2\text{O}_4)$. N_2O_5 .

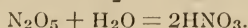
The first, second and fifth act as anhydrides, forming hyp-nitrous, nitrous, nitric acids. The third and fourth form a mixture of HNO_2 and HNO_3 . All the oxides are prepared from HNO_3 .

Nitric Anhydride N_2O_5 —

Discovered by Deville. Made by passing Cl over dry AgNO_3 —apparatus entirely of glass, melted together.



White, unstable solid—with H_2O forms HNO_3 .



NITRIC ACID— HNO_3 .

History—

Known early as 9th century when it was made by distillation of $\text{ZnSO}_4 + \text{KNO}_3$ and called “Aqua fortis,” “Parting water,” or “Spiritus nitri fumans Glauberi.”

Occurrence—

Not free—in salts as nitrates, is widely distributed in the earth especially in Chili as NaNO_3 or Chili Saltpeter—in small quantities in air and water.

Formation—

From elements by electric spark in presence of H_2O —also in process of decay, occasioned by bacteria, the so-called “nitrification,” which takes place best in the dark and is stopped by destruction of the bacteria through chloroform or boiling.

Preparation—

Action of H_2SO_4 on K salt.



Impurities—

Ordinary KNO_3 contains KCl ; hence, HCl is formed with crude HNO_3 —another common impurity is H_2SO_4 .

Properties—

Pure HNO_3 is a colorless volatile liquid leaving no residue—no stronger acid has been yet obtained—so powerful an oxidizing agent that it deflagrates with easily combustible substances—exposed to sunlight it gradually decomposes to its oxides and H_2O ; hence, strong acid is often colored by NO_2 —a powerful mono-basic acid; it dissolves most of the metals forming nitrates all of which are soluble in water, some insoluble in HNO_3 , all decompose at high heat.

Used—

Commercially to etch copper.

Aqua Regia—

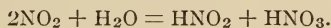
Formed by mixing $\text{HNO}_3 + \text{HCl}$; acts as nascent chlorine, forming a chloride.



A long known mixture of dark yellow color and suffocating fumes and odor—called “royal water” because it dissolves “noble” metals, gold and platinum.

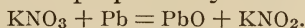
Nitrogen Peroxide NO_2 —

Whenever NO is generated it combines with O of air to NO_2 —red brown gas—poisonous, suffocating odor—very corrosive—colors organic bodies yellow—strong oxidizing agent, C and P burn in its vapor—easily liquified and solidified. NO_2 does not form a hydrate, but with H_2O gives



Nitrous acid HNO_2 —

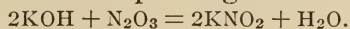
Hydrate of N_2O_3 and formed by heating HNO_3 with starch, sugar, or other easily oxidizable substance, which reduces HNO_3 . In same way nitrites are prepared by reduction of nitrates—



Nitrites are mostly soluble in H_2O .

Nitrous anhydride N_2O_3 —

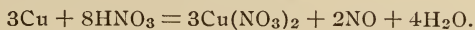
Formed by decomposition of nitrites—a red brown gas easily condensed to indigo blue liquid. Passing the gas into alkali hydrates forms corresponding nitrites—



Nitric Oxide NO —

Never found free, as it combines with O of air to form NO_2 .

Made by action of HNO_3 on Cu .



At high temperature N_2O is formed.

Colorless gas, irrespirable, poisonous—oxidizes in air to NO_2 —slightly soluble in H_2O . NO supports combustion of P and S but not of ordinary substances, as wood.

Hyponitrous Acid HNO —

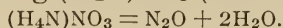
Discovered in 1871 by reducing KNO_3 with NaHg . Known only as salt.

Nitrous oxide N_2O —

Discovered by Priestley 1776.

Colorless gas—sweet taste—will not support life, but can be breathed a short time, producing intoxication and anæsthesia (laughing gas). N_2O supports combustion like O but extinguishes burning S. Condenses easily and forms with ether $(C_2H_5)_2O$ the coldest known mixture—distinguished from O by leaving a residue of N when exploded with H.

Prepared by heating $(H_4N)NO_3$ (ammonium nitrate)—



At high temperature NO forms with explosion.

Valence

is that property of an element by which it combines with definite masses of other elements. Unit of valence is hydrogen. Elements are classified according to their power of uniting with or replacing different proportions of the unit. Thus chlorine unites with one H to form HCl and is a univalent element. Oxygen forms H_2O , requiring two parts of H and is a bivalent element. Nitrogen forms H_3N , acting as a trivalent element. Carbon forms H_4C , acting as a quadrivalent element. In a comparatively few instances valence requires five, six or seven hydrogens. Under the same conditions, valence of a given element is always constant, but it may vary for the same substance in different compounds, depending on what element it unites with and under what conditions union takes place. Generally when a substance is present in smaller mass it unites with higher valence. Thus a small amount of carbon heated in presence of much oxygen forms CO_2 , where C is quadrivalent, but where a large amount of C is heated with small O the gas formed is CO, where C acts as a bivalent element. The greatest variation appears in the oxides and chlorides. The oxides of N for example— N_2O , N_2O_2 , N_2O_3 , N_2O_4 , N_2O_5 —give to nitrogen five distinct valences; and other elements form similar series.

In general, valence may be considered less as a property of the specific elements and more as a function incident to their combination.

PHOSPHORUS—P.

At. Wt. 31. Val. III or V.

History—

Discovered in 1667 by Brand when searching for philosopher's stone in urine—which was the only source till 1750—later Scheele obtained it from bones.

Occurrence—

In soil as apatite and pyromorphite—extracted from soil by plants—then to animals in bones, brain and urine.

Properties—

Occurs in two important allotropic forms, yellow crystalline and red amorphous phosphorus. Yellow is converted by heat (250°) to red—ordinary P exposed to light becomes red on surface—common or yellow P is a wax-like solid—ozone odor—luminous in the dark—oxidizes easily—burns at 44° —a mixture of P and KClO_3 detonates when struck—causes chronic poisoning—almost insoluble in water, but dissolves easily in CS_2 —preserved in H_2O . Red P is a weaker chemical agent—odorless, non-phosphorescent—not poisonous—does not dissolve in CS_2 —is stable at ordinary temperatures—at 260°C red changes back to yellow.

Preparation—

Bones are burned to remove organic matter (55% whole). Ash treated with H_2SO_4 —this changes insoluble $\text{Ca}_3(\text{PO}_4)_2$ of the bones to a soluble phosphate which heated with charcoal gives free P—the essential reaction with C is $2\text{P}_2\text{O}_5 + 5\text{C} = \text{P}_4 + 5\text{CO}_2$. P is purified by redistillation. Cannot be successfully prepared on small scale.

Use—

Chiefly for making matches. Safety matches = mixture of potassium-chlorate (KClO_3) and antimony trisulphide (Sb_2S_3)—the surface = red phosphorus + MnO_2 . Ordinary match = yellow P, some oxidizing agent as $\text{MnO}_2 + \text{KClO}_3$ and glue.

PHOSPHINE— PH_3 .

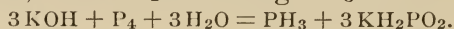
Properties—

The only important compound of P and H—a gas of disa-

greeable odor and highly poisonous—chemical action similar to NH_3 —forms phosphonium compounds analogous to H_4N compounds. P and H form also P_2H_4 , a liquid, and P_4H_2 , a gas, but these are unimportant. PH_3 does not ignite spontaneously, except in presence of P_2H_4 .

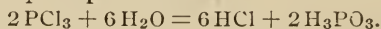
Preparation—

From KOH , P and H_2O forming PH_3 and an acid of P.



Phosphorus tri-chloride PCl_3 —

The more important chloride of P—prepared by action of Cl on melted P—a colorless liquid which fumes in moist air and with water forms phosphorus acid.



There is also a penta-chloride prepared in same way but with excess of Cl—a white crystalline substance.—I and Br form similar compounds, all of which are much used in the arts.

Phosphorus Oxides—

P forms two oxides by burning in greater or less quantity of dry air, P_2O_3 and P_2O_5 —both white feathery solids— P_2O_5 is more important and generally called phosphoric anhydride—unites eagerly with water in different proportions to form acids.

- (1) $\text{P}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HPO}_3$ (Meta- or glacial phosphoric acid).
- (2) $\text{P}_2\text{O}_5 + 2\text{H}_2\text{O} = \text{H}_4\text{P}_2\text{O}_7$ (Pyrophosphoric acid).
- (3) $\text{P}_2\text{O}_5 + 3\text{H}_2\text{O} = 2\text{H}_3\text{PO}_4$ (Orthophosphoric acid).

Orthophosphoric acid H_3PO_4 —

P is oxidized with HNO_3 and residue evaporated—difficult operation as oxidation is slow and does not work with weak while it explodes with strong HNO_3 —can be made also from the Ca salt but is difficult to purify.

Salts—

Form three classes— M_2PO_4 , M_2HPO_4 , M_3PO_4 . Most stable are the alkaline M_2HPO_4 or of heavy metals M_3PO_4 .

By ignition, M_3PO_4 unchanged.

By ignition, M_2HPO_4 changed to pyrophosphate $\text{M}_4\text{P}_2\text{O}_7$.

By ignition, MH_2PO_4 changed to metaphosphate MPO_3 .

Pyrophosphoric Acid $\text{H}_4\text{P}_2\text{O}_7$ —

Prepared by heating H_3PO_4 or adding H_2O to P_2O_5 as above.

Gradually absorbs H_2O forming ortho- acid.

$\text{H}_4\text{P}_2\text{O}_7$ is a tetrabasic acid but acts as dibasic, forming $\text{M}_4\text{P}_2\text{O}_7$ and $\text{M}_2\text{H}_2\text{P}_2\text{O}_7$.

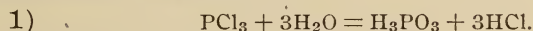
All normal salts are soluble save those of alkalis.

All ortho- salts change with heat to pyro- salts.

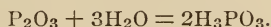
Metaphosphoric Acid HPO_3 —

Ortho- acid slightly ignited gives pyro- and strongly gives meta—the melted mass solidifies to “glacial” meta- acid—this dissolves in H_2O and acts like P_2O_5 in H_2O —solution coagulates albumen and gives white precipitate with AgNO_3 or BaCl_2 —it is gradually changed to ortho-.

Phosphorous Acid H_3PO_3 —



2) Slow oxidation of P in moist air.

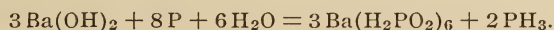


Strong reducing agent, oxidizing to H_3PO_4 .

Forms two salts, MH_2PO_3 and M_2HPO_3 .

Hypophosphorous Acid H_3PO_2 —

Anhydrite P_2O has not been obtained—free acid formed by treating Ba salt with H_2SO_4 —Ba salt formed in same reaction as for PH_3 .



White crystalline mass—a monobasic acid—Ca salt used in medicine.

ARSENIC—As.

At. Wt. 75. Val. I, III or V.

History—

Element long known.

Occurrence—

Widely distributed as native As and compounds, especially sulphides As_2S_2 realgar, As_2S_3 orpiment and mispickel FeAsS —As known in commerce as “cobalt” or “flystone.”

Properties—

Two allotropic modification—amorphous black powder, and the common crystalline form—a brittle, steel grey solid with metallic lustre—vapor is yellow with garlic odor—burns with blue flame to As_2O_3 —metallic As is not poisonous.

Preparation—

- (1) Heating mispickel in clay cylinders.
- (2) Heating As_2O_3 with charcoal and resubliming the product.

Use—

Pigments—flypaper—also forms alloys as with Pb to make shot.

HYDROGEN ARSENIDE, ARSINE— AsH_3 .

History—

Discovered by Scheele.

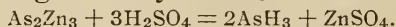
Properties—

Colorless gas—disagreeable odor—violent poison—burns with blue flame to $\text{As}_2\text{O}_3 + \text{H}_2\text{O}$ —with limited supply of air to $\text{As} + \text{H}_2\text{O}$ —Decomposed at red heat to $\text{H} + \text{As}$ which is deposited as a mirror giving a delicate test for As.

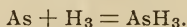
As in reagents— FeS_2 usually contains some arsenic, whence it passes into H_2SO_4 , HCl , and any reagents made from these acids.

Preparation—

- (1) By heating the alloy As_2Zn_3 with an acid.



- (2) Formed when As is treated with nascent H.

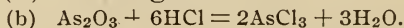


- (3) When compounds of As are treated with organic bodies especially decaying bodies, thus AsH_3 is formed by As in wall paper.

Arsenic trichloride (AsCl_3)—

Formed by

- (a) burning As in Cl.



—Heavy, colorless, fuming liquid—important from its volatility as it may cause loss of As in analysis.

As—OXIDES—ACIDS AND SULPHIDES.

Oxides—

The oxides As_2O_3 and As_2O_5 correspond to those of P and N and are of acid nature— As_2O_5 combines directly with water, forming H_3AsO_4 , while As_2O_3 does not easily dissolve; hence is known as *arsenious acid*.

Arsenic trioxide (As_2O_3)—

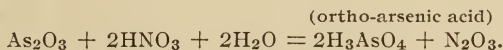
The common compound of As—commercially known as “arsenic” or “white arsenic”—a white amorphous powder—formed when As burns—slightly soluble in H_2O with faint acid reaction—odorless—sweet metallic taste—deadly poison— As_2O_3 unites with bases, forming *arsenites*— H_2S precipitates As_2S_3 , which is soluble in $(\text{H}_4\text{N})_2\text{S}$.

Arsenic pentoxide (As_2O_5)—

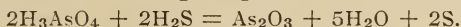
A colorless deliquescent mass—formed when As_2O_3 is heated to a red heat—not when As burns—its water solution forms arsenic acid (H_3AsO_4 .)

Arsenic Acid (H_3AsO_4)—

White crystalline solid, formed when As_2O_3 is digested in HNO_3 .

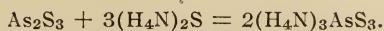


If heat is applied the pyro- and metarsenic acids are formed corresponding in formula to the phosphoric series— $\text{Cu}_3(\text{AsO}_4)_2$ is blue— Ag_3AsO_4 brown—with H_2S arsenic acid is generally first reduced to As_2O_3 then precipitated as As_2S_3 .



ARSENIC SULPHIDE.

Chief is Orpiment As_2S_3 —volatile without decomposition—heated with reducing agent to As—soluble in alkaline sulphides, forming salts—



A less important sulphide is Realgar (As_2S_2).

ANTIMONY—Sb.

At. Wt. 120. Val. I, III, or V.

History—

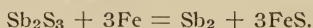
One of old metals—called Stibium by Pliny.

Occurrence—

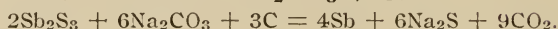
Occurs native and in combination—common ore is stibnite (Sb_2S_3).

Preparation—

(1) Heat stibnite with Fe.



(2) Heat stibnite with $\text{Na}_2\text{CO}_3 + \text{C}$.



Properties—

Silver white metal, highly crystalline, hard and brittle—volatilizes at red heat—expands on cooling—oxidizes slowly in moist air, rapidly in HNO_3 to Sb_2O_3 —Precipitates from its solution black by Zn—soluble in strong HCl and $\text{H}_2(\text{C}_4\text{H}_4\text{O}_6)^{\text{(tartaric acid)}}$ —insoluble in HNO_3 —Sb compounds are poisons.

Used—

By reason of its expansion on cooling Sb is used in type metal.

Stibine (H_3Sb)—

Analogous to H_3N and H_3As —discovered by Marsh process—formed under same conditions as AsH_3 , and has same properties—chiefly important in detection of Sb.

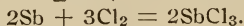
ANTIMONY HALOIDS.

Sb forms two compounds with Cl and F— SbCl_3 , SbCl_5 — SbF_3 , SbF_5 —with Br and I it forms SbBr_3 and SbI_3 .

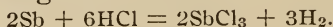
Antimony trichloride (SbCl_3)—

A soft colorless body—commercially called “Butter of Antimony”—prepared:

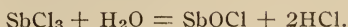
(1) By heating in Cl gas (note PCl_3).



(2) By dissolving Sb in HCl .



SbCl_3 is decomposed by H_2O with precipitation of the oxychloride or basic chloride SbOCl .



This white precipitate is a characteristic reaction for antimony— SbOCl is called “Powder of Algaroth.”

Antimony pentachloride (SbCl_5)—

Prepared by action of chlorine in excess on SbCl_3 , and purified by distilling in current of Cl — SbCl_5 is a fuming colorless liquid, decomposed by water to SbO_2Cl . Other haloid compounds of Sb are analogous to those of P and As.

OXIDES OF Sb.

Three compounds— Sb_2O_3 , Sb_2O_4 , Sb_2O_5 .

Sb_2O_3 acts as base with strong acids.

Sb_2O_4 base with strong acid—acid with strong base.

Sb_2O_5 acts as acid.

Antimony trioxide (Sb_2O_3)—

Found free in prismatic crystals.

Prepared:

- (1) Heating Sb in the air— Sb_2O_3 sublimes as white crystals.
- (2) Oxidizing Sb with HNO_3 .
- (3) “ “ “ KNO_3 .

Sb_2O_3 is isomorphous with As_2O_3 —acts as a base—soluble in strong HCl , forming SbCl_3 , which with H_2O gives SbOCl .—ANTIMONIC ANHYDRIDE Sb_2O_5 is a pale yellow powder which reduces to Sb_2O_3 with heat.

Antimonic Acid H_3SbO_4 —

Analogous to H_3AsO_4 and H_3PO_4 —derived from Sb_2O_5 . The Na salt is insoluble in H_2O and of importance in separating As from Sb.

Antimony Sulphides—

Two known sulphides, Sb_2S_3 and Sb_2S_5 , analogous to sulphides of As. The trisulphide, stibnite, Sb_2S_3 is most important ore of antimony—separated from impurities by melting—crude Sb_2S_3 called “antimony” oxidizes with heat to Sb_2O_3 and Sb_2O_4 . Antimony pentasulphite Sb_2S_5 is an orange colored powder formed by passing H_2S into antimonic acid.

BISMUTH—Bi.

At. Wt. 210. Val. I, III, or V.

Occurrence—

Mostly as *free* Bi in Saxony also as Bi_2O_3 and Bi_2S_3 —comparatively rare substance.

Properties—

Brittle, hard, faint red-white color, with metallic lustre—not easily oxidized but at red heat converted to Bi_2O_3 —not attacked by dilute H_2SO_4 or HCl , but easily dissolved by HNO_3 —melting point lowered by alloying with other substances, forming fusible metals, as Woods metal which melts in hot water—Bi does not combine with H—unites with Cl and Bi, in two proportions—with I in one.

BISMUTH CHLORIDES.

Bi forms two chlorides, Bi_2Cl_4 and BiCl_3 .

Bismuthous Chloride (Bi_2Cl_4) is a black, water-absorbing powder, formed by heating Bi and Hg_2Cl_2 .



Bismuth Chloride (BiCl_3) is volatile like SbCl_3 and formed in same way, also by dissolving the oxide in HCl —a soft unstable mass, decomposed by H_2O to BiOCl .

Bismuth oxide (Bi_2O_3)—

Formed by

- (1) Heating Bi in the air.
- (2) Heating BiONO_3 .

A yellow powder—forms salts with acids which are colorless when the acid is colorless—All salts decomposed by much H_2O —chief compound is the basic nitrate BiONO_3 , BISMUTH SUB-NITRATE, or bismuthyl nitrate, which is much used in medicine, and as cosmetic. Often contains a dangerous amount of As.

BORON—B.

At. Wt. 11. Val. III.

History—

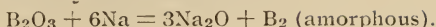
Sodium salt long known as “Tinkal.”

Occurrence—

Not free—combined as boracic acid H_3BO_3 , and borax $\text{Na}_2\text{B}_4\text{O}_7$.

Preparation—

Treat Boric anhydride with Na.



Properties—

Two varieties, amorphous and crystalline.

Amorphous is a brown powder—odorless, tasteless, insoluble in H_2O , non-conductor of electricity—on heating combines directly with many substances as S, Cl, Br—BORON burns in air to B_2O_3 and N.

Crystalline variety obtained by igniting B with Cl—thus formed is hard as the diamond.

BORACIC ACID— H_3BO_3 .

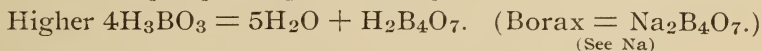
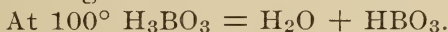
Occurrence—

In volcanic districts of Tuscany certain hot springs rise, called “fumaroles”—these carry H_3BO_3 mechanically. Steam from the fumaroles is passed into water and 1% solution of H_3BO_3 obtained—this is evaporated by heat of the springs and purified through its sodium salt.

Properties—

Acid and its salts form, on heating, a colorless glass bead, which dissolves oxides with characteristic colors.

H_3BO_3 itself forms no salts but on heating it loses water, condensing to other acids which form salts.



(See Na)



Borates are identified by a green color to the flame. Alkali borates only are soluble in water.

Boron trioxide (B_2O_3)—

The only oxide of Boron—called Boracic anhydride. $3\text{H}_2\text{O} + \text{B}_2\text{O}_3 = 2\text{H}_3\text{BO}_3$. B_2O_3 is the source of all B compounds, and is prepared by heating B in the air or heating H_3BO_3 .



Boron Haloids—

B unites with all but I—forms the tri-compounds BiCl_3 , etc. —Liquids of low b.p. and easily decomposed by H_2O —formed by passing the haloid over heated B or $\text{B}_2\text{O}_3 + \text{C}$.

SILICON—Si.

At. Wt. 28. Val. II or IV.

History—

Isolated by Berzelius in 1823. Amorphous variety found first.

Occurrence—

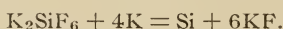
Always combined with oxygen in Silicon dioxide (SiO_2) as quartz, sand, etc.—All geological formations except chalk contain Si.

Preparation—

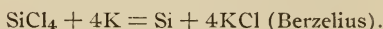
Three varieties of Silicon.

(1) Amorphous, made by

(a) fusing potassium-fluo-silicate K_2SiF_6 with K or NA.



(b) use Silicon tetra-chloride and K.



(2) Graphitic, by fusing the amorphous with Al.

(3) Crystalline, by fusing the graphitic.

Properties—

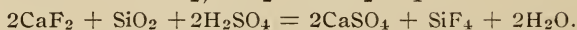
Si has strong affinity for O—the amorphous variety burns to SiO_2 —Si is soluble in Aluminum (Al) just as C is in iron. Also soluble in HF forming $\text{SiF}_4 + \text{H}_4$.

Silicic Hydride (SiH_4)—

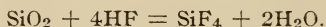
A spontaneously lighting gas—burns to SiO_2 and H_2O —analogous to methane (CH_4).

Silicon Fluoride (SiF_4).

Prepared from CaF_2 , SiO_2 and H_2SO_4 .



Etching of glass depends on this reaction as the SiO_2 may be in form of glass.



Properties—

Colorless gas—bad odor—deadly poison—fumes with H_2O , forming silicic acid and hydro-fluo-silicic acid.



Hydro-Fluo-Silicic Acid (H_2SiF_6)—

Prepared by passing SiF_4 into H_2O .

K salt is insoluble in H_2O .

Chemical deportment similar to hydrogen-halogen acids (HCl , etc).

Silicon Dioxide SiO_2 —"Silica"—

At first supposed to be an earth—called vitreous earth from use in glassmaking—finally shown to be a weak acid.

Found both free and combined in nature. Free as quartz (anhydrous), opal (hydrous), amethyst quartz (colored with Mn), agate, flint, etc. In combination as silicates forming most of the "rocks".

By evaporating to dryness with acids, silicates decompose with a residue of SiO_2 insoluble in acids; this constitutes a *test* for silica. Silicates are insoluble save the Na or K salt.

Silicic Acid H_2SiO_3 —

There are many silicic acids but this is the most common form—Heated it loses water to form SiO_2 .

CARBON—C.

At. Wt. 12. Val. II. or IV.

History—

First referred to as an element by Lavoisier 1780. He showed that "mephitic air" or "carbonic acid" contained C and O, also that C was the important part of charcoal.

Occurrence—

Free in three modifications—crystalline as diamond and graphite—amorphous as various kinds of coal—as CO_2 in air and water—as carbonates in all soil, sometimes forming mountains. C crystallizes easily as graphite, but not easily as diamond. Obtained as amorphous C from organic bodies.

Diamond—

The diamond is found mostly in old rock strata chiefly in India, Borneo, Brazil and South Africa. Usually colorless—when crude covered with dull crust—when polished exhibits great lustre—high refraction—extreme hardness—brittle—poor conductor of heat and electricity—not easily attacked by oxi-

dizing agents—unchanged by heat alone—burns easily in oxygen blowpipe.

Carbonado or black diamonds occur in large masses—have hardness but not fire or water.

Graphite—

Widely distributed as a mineral—called also plumbago and “black lead”—Till 1798 thought to contain Pb—usually contains about five % ash ($\text{SiO}_2 + \text{Fe}_2\text{O}_3$).

A black substance—metallic lustre—very friable but exceedingly hard—Good conductor of heat and electricity—Crystallizes in hexagonal crystals from melted Fe—not affected by O at ordinary temperatures—at high temperature oxidizes to CO_2 —insoluble in all ordinary liquids.

USED for lead pencils—lubricants—electrotyping—as preservative from rust—mixed with clay is used for crucibles.

CHARCOAL.

Preparation—

Not easily obtained pure—made by distilling wood in retorts or by burning wood in small amount of air, getting rid of volatile substances and saving most of the carbon.

Properties—

Charcoal is brittle because of cell structure of wood which is retained.

Physical properties are important—absorbs and condenses great variety of gases and vapors, especially such as dissolve in H_2O —action is not mere absorption, but oxidation of offensive gases, hence charcoal is a valuable disinfectant, by hastening destruction of putrescible organic matter—is not an antiseptic or preservative agent—C likewise decolorizes, being especially used in purification of sugar, and in chemical and pharmaceutical preparations.

Bone Black—

Prepared by destructive distillation of bones—particularly good for all uses of charcoal because of large surface.

Coke—

By destructive distillation of soft coal—chemically classed between graphite and charcoal.

Lamp-black—

The kind of charcoal, from burning oils in small amount of air.

Gas Carbon—

Carbon of gas retorts is formed by long heating of the crust which condenses on the interior of gas retorts—very hard, compact and dense—conducts heat and electricity—employed in galvanic batteries and as pencils for electric lamps. Chemically classed between charcoal and graphite.

Coal—

Chief varieties are Anthracite, Bituminous and Peat.

The amount of carbon decreases while water and S increase in the given order.

CARBON AND HYDROGEN.

Form an immense number of compounds known as hydrocarbons. One is formed by direct union of elements— C_2H_2 , ACETYLENE—formed with C electrodes in atmosphere of H.

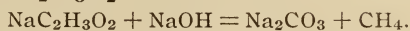
METHANE— CH_4 (MARSH GAS).

Occurrence—

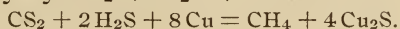
Found where vegetable matter decays under water—hence called Marsh gas—Found also in coal mines where it often causes explosions, hence named “FIRE DAMP”—chief constituent of gas wells.

Preparation—

Distilling $NaC_2H_3O_2$ with $NaOH$ —



Synthetically by $CS_2 + H_2S +$ red hot Cu —



Properties—

Colorless, inodorous, tasteless gas (condensed by 188 at.)—slightly soluble in water—burns with slight luminous flame—mixture with air explodes (“fire damp”), forming $H_2O + CO_2$, “CHOKE DAMP.”

CARBON MONOXIDE— CO .

Properties—

A colorless gas, inodorous, insoluble in H_2O —burns with

pale blue flame to CO_2 —highly poisonous—a powerful reducing agent, forming CO_2 .

Preparation—

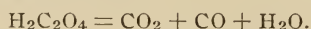
Always formed by incomplete combustion of C. At first prepared by distilling ZnO with C and supposed to contain H.

a) Pass steam over red-hot coal.



This is “water gas,” one form of illuminating gas—the other or coal gas, formed by destructive distillation of coal, is less poisonous.

b) CO is formed by decomposition of many organic substances, especially oxalic acid.



CARBON DIOXIDE— CO_2 .

History—

Known in earliest times through action of vinegar on chalk—first distinguished as a gas by V. Helmont, who called it “gas sylvestre”—Lavoiser termed it carbonic acid and showed its composition.

Occurrence—

Widely distributed in atmosphere and soil—occurs in liquid form in some crystalline substances—generally found united with bases, especially in CaCO_3 .

Preparation—

By action of acids on carbonates—



Properties—

Colorless gas—acid odor—heavy—incombustible—extinguishes flame—will not support respiration—liquid form is colorless—becomes solid on rapid evaporation, and forms with ether a freezing mixture— CO_2 is absorbed most by water at low temperature, making carbonic acid waters—its presence in wine and beer is due to fermentation—true carbonic acid is H_2CO_3 , a very unstable substance forming two classes of salts—carbonates of alkalies only are stable, but acid salts of alkalies are decomposed by heat.



CARBON DISULPHIDE—CS₂.

History—

Discovered 1796 by Lampadius and called "sulphur alcohol." Thought to contain S, H, C, N, but in 1811 shown to be only C and S.

Preparation—

Pass vapor of S over red-hot C.

Properties—

Heavy, colorless, strongly refracting liquid—etherial odor when fresh—volatile, with poisonous vapor, having cumulative effect—easily ignited and burns to CO₂ + SO₂—vapor burns with NO—mixed with air its vapor explodes—remarkable antiseptic power—Decomposed by heat to its elements—CS₂ forms salts of thiocarbonic acid H₂CS₃.

CS₂ used chiefly as solvent especially of S and P, also of Br and I.

CYANOGEN—C₂N₂.

History—

Discovered in 1815 by Gay-Lussac while investigating "prussic acid"—called it cyanogen from prussian blue.

Preparation—

C and N unite only with difficulty—the compounds are obtained indirectly and all contain the group CN, which was the first "radicle" discovered.

1) C₂N₂ formed when organic bodies are heated with K or Na—detection of N in organic bodies depends on this.

2) Heat Hg(CN)₂ = Hg + C₂N₂.

Properties—

Colorless gas—strong odor—intensely poisonous—burns with red flame to CO₂ + N; CN is recognized by its formation of Prussian blue—when in free state the radicle (CN) exists as C₂N₂ DICYANOGEN, which closely resembles the Haloids.

HYDROCYANIC ACID—HCN

Occurs in nature in kernels of almonds, peaches, plums, etc.; blossoms and leaves of the peach tree, and several other trees

and shrubs—a liquid similar in properties to HCl, but intensely poisonous—generally prepared by treating K salt with HCl,



ACETIC ACID— $\text{H}(\text{C}_2\text{H}_3\text{O}_2)$

Pyroligneous acid a bi-product in making charcoal—clear, colorless liquid—characteristic taste and odor. Pure acid is called “glacial”—dilute, impure form is known as vinegar— $\text{H}(\text{C}_2\text{H}_3\text{O}_2)$ is a monobasic acid and forms acetates $\text{M}(\text{C}_2\text{H}_3\text{O}_2)$.

OXALIC ACID— $\text{H}_2\text{C}_2\text{O}_4$.

Crystalline white solid—poisonous—good reducing agent. $\text{H}_2\text{C}_2\text{O}_4$ is a dibasic acid formed by action of HNO_3 on many organic substances, especially the sugars, starch, etc.—wholesale by action of KOH on sawdust—occurs widely distributed, especially in plants of oxalis variety and in urinary calculi.

CHROMIUM—Cr.

At. Wt. 52.4. Val. II or III.

History—

In 1797 was discovered by Vanquelin in mineral “crocoisite” (PbCrO_4).

Occurrence—

Not widely distributed—chief ore is “chromite” or chrome iron ore ($\text{FeO} \cdot \text{Cr}_2\text{O}_3$).

Extraction—

Ore is pulverized and heated with K_2CO_3 and CaO — K_2CrO_4 POTASSIUM CHROMATE is formed and extracted with water— H_2SO_4 is added to form $\text{K}_2\text{Cr}_2\text{O}_7$ POTASSIUM BICHROMATE, which is the commercial salt. The METALLIC Cr is obtained by heating Cr_2O_3 (chromic oxide) with C.

Properties—

A hard, difficultly fusible metal—similar to iron—with heat it slowly oxidizes to Cr_2O_3 —burns in oxygen with bright light—soluble in HCl and H_2SO_4 —not altered by HNO_3 —forms chromous and chromic compounds—Cr forms an immense number of compounds, in some it is acidic, in some basic, or its acid and basic oxides may unite, forming a salt of itself.

CHROMIUM HALOIDS— CrCl_2 , Cr_2Cl_6 , CrF_6 .

Chromous chloride—

A white powder soluble in water—formed by reducing Cr_2Cl_6 with H—chromous compounds are hard to obtain and easily oxidize to chromic.

Chromic Chloride Cr_2Cl_6 —

A violet sublimate obtained by ignition of Cr_2O_3 with C in a current of Cl—insoluble when pure but easily dissolves if a trace of CrCl_2 is present—the hydrated salt $\text{Cr}_2\text{Cl}_2 \cdot 12\text{H}_2\text{O}$ forms green deliquescent crystals.

Chromic Fluoride CrF_6 —

Chiefly interesting as showing the sexivalent character of Cr.

CHROMIUM OXIDES.

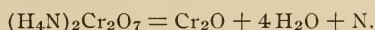
CrO — Cr_2O_3 — CrO_3 —first two basic—last acidic.

Chromous Oxide CrO —

Known only in its HYDRATE $\text{Cr}(\text{OH})_2$, which is made from $\text{CrCl}_2 + \text{KOH}$ and rapidly changes to chromic oxide.

Chromic Oxide Cr_2O_3 —

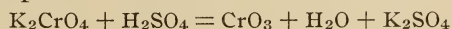
Found in nature as “chromite”—formed as a green powder by ignition of ammonium di-chromate



The ignited oxide is insoluble in acids—fused with silicates it gives them an emerald green color—used to color glass and porcelain—when freshly precipitated the hydrate is soluble in acids. It is used as a paint, “Gingnet’s green”—like all sesquioxides, Cr_2O_3 is normally basic, but will not give salts with weak acids. On the contrary, with strong bases it exhibits an acid character, forming chromites.

Chromic anhydride CrO_3 —

Red deliquescent crystals—active oxidizing agent—highly corrosive—made from the K chromate or bichromate by treating with H_2SO_4 —



The water solution contains the unstable chromic acid, H_2CrO_4 , which is analogous to H_2SO_4 but known only through its salts the chromates and bichromates—the neutral chromates are gen-

erally yellow, the bichromates red—basic Cr_2O_3 unites with acidic CrO_3 to form chromium chromate, Cr_2CrO_4 —The Ba, Pb, Ag, and Hg salts of chromic acid are insoluble.

Chrome alum $\text{Cr}_2 \cdot \text{K}_2(\text{SO}_4)_4 + 24 \text{H}_2\text{O}$ —

The chief salt where Cr acts as a base—formed as violet octohedra, when a solution of $\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{SO}_4$ is treated with SO_2 , and evaporated below 80° —above 80° the solution turns green and refuses to crystallize.

Chromium sulphides CrS and Cr_2S_3 —

Cr forms two sulphides, CrS and Cr_2S_3 corresponding to the oxides but cannot be formed in the wet way.

MOLYBDENUM—MO.

At. Wt. 96. Val. II, IV, or VI.

History—

Name comes from the Greek for graphite, with which the mineral Molybdenite was confused— MoO_3 was separated in 1778 by Scheele and Mo in 1798 by Hjelm.

Occurrence—

Chief ore is molybdenite MoS_2 .

Properties—

Hard, silver white metal—very stable—when pure, infusible—after long heating converted to Mo_2O_3 —dissolves in HNO_3 and aqua regia.

COMPOUNDS OF Mo.

Chief compounds are the acid H_2MoO_4 and its (NH_4) salt $(\text{NH}_4)_2\text{MoO}_4$.

Molybdic Oxide MoO_3 —

The anhydride of molybdic acid is MoO_3 , made by roasting molybdenite—it combines readily with bases, forming molybdates.

COMPOUNDS OF Mo.

Molybdic acid H_2MoO_4 —

Is formed when molybdates are treated with dilute HNO_3 —Dissolved in strong ammonia it forms $(\text{H}_4\text{N})_2\text{MoO}_4$, AMMONIUM MOLYBDATE, which in HNO_3 solution is used to precipitate

H_3PO_4 in the form of ammonio-phospho-molybdate, a delicate TEST for PHOSPHATES.

TUNGSTEN—W.

At. Wt. 184. Val. II, IV, or VI.

Analogous to Mo—forms similar compounds—Mo compounds crystallize easier—Tungstic acid similar to molybdic acid—common salts are Na and H_4N tungstates—the former used to make fabrics fire-proof—chief ore is wolframite, FeWO_4 . CALCIUM TUNGSTATE, CaWO_4 , is used as a fluorescent screen.

TIN—Sn.

At. Wt. 118. Val. II or IV.

History—

Used by the Phœnicians, who got it from England.

Occurrence—

Native in small amounts—chief ore is tin stone SnO_2 .

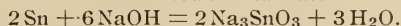
Extraction—

By reducing the ore with carbon and remelting.

Properties—

Lustrous white metal—crystalline—when bent emits a peculiar sound (tin cry)—ductile—malleable, but more so at 100° —brittle at 200° —not oxidized by the air till its melting point—at white heat burns brilliantly to SnO_2 —forms salts with HCl and H_2SO_4 —forms SnO_2 with HNO_3 or with the alkalies, when it acts as an acid.

sodium stannate



Sn is not affected by H_2S .

Use—

Chiefly in “tinware” which is iron covered with a layer of tin—also in bronze (Cu, Sn, Zn), solder (Sn, Pb), Britannia metal (Sn, Sb), tin amalgam (Sn, Hg)—the last is used in silvering mirrors.

TIN CHLORIDES.

Stannous chloride SnCl_2 —

The anhydrous form is made by HCl gas on metallic tin—hydrated by treating excess of Sn with HCl — SnCl_2 comes on the

market as TIN SALT—is a white crystalline substance—soluble in small amount of H_2O , in large amount forms the basic chloride—has a strong tendency to unite with Cl , hence used as a reducing agent to form $SnCl_4$.

Stannic Chloride $SnCl_4$ —

A colorless, fuming liquid, hence name of “*Spiritus fumans Libavii*”—made by action of Cl on Sn or $SnCl_2$ —combines with metallic chlorides to form double salts, of which “pink salt” $SnCl_4 \cdot 2H_4NCl$ is most important—both chlorides of tin are used as mordants—the bromides and iodides are analogous.

TIN OXIDES.

Stannous Oxide SnO —

The basis of stannous salts—formed by heating the hydrate— SnO is a white powder which easily oxides to SnO_2 .

Stannic Oxide SnO_2 —

“Tinstone” is a crystalline solid formed by heating Sn or SnO —insoluble in acids—fused with $NaOH$ or KOH it forms soluble stannates— SnO_2 acts also as a weak base forming with H_2SO_4 stannic sulphate $Sn(SO_4)_2$

ACIDS OF TIN.

Stannic Acid H_2SnO_3 —

Separates as a white precipitate when HCl is added to potassium stannate (K_2SnO_3)—dissolves easily in HNO_3 , HCl and the alkalies—left under water or in vacuo, stannic gradually changes to METASTANNIC acid, an insoluble modification of the same composition. This isomer is also formed when tin is treated with HNO_3 , concentrated—metastannates are entirely different from stannates.

Stannic Phosphate $Sn_3(PO_4)_2$ —

Tin forms with phosphoric acid a compound, $Sn_3(PO_4)_2$, which is insoluble in HNO_3 —used in analysis to remove H_3PO_4 .

TIN SULPHIDES.

Stannous Sulphide SnS —

Formed by direct union at high temperature—also precipitated from stannous salts by H_2S in brown amorphous form.

Stannic Sulphide SnS_2 —

Precipitated as a yellow powder, by H_2S , from stannic salts—when sublimed it forms a bright yellow crystalline mass known as mosaic gold which is used in bronzing.

STANNOUS SALTS.

If acid is colorless the salt is either colorless or yellow—stannous salts have metallic taste—absorb O from the atmosphere—change easily to stannic—

KOH gives white $\text{Sn}(\text{OH})_2$ soluble in excess.

$\text{H}_4\text{N}(\text{OH})$ gives white $\text{Sn}(\text{OH})_2$ insoluble in excess.

AuCl_3 gives purple of Cassius, characteristic.

HgCl_2 gives Hg_2Cl_2 , characteristic.

TITANIUM Ti —ZIRCONIUM Zr —THORIUM Th .

At. Wt. 48. Val. IV. At. Wt. 91. Val. IV. At. Wt. 232. Val. IV.

Three rare metals closely resembling tin but more basic in character and forming NO-ous salts—The most important is Ti which generally accompanies iron, especially in Titanic Iron, FeTiO_3 —metal is obtained by decomposing potassium fluo-titanate, K_2TiFe , with K (see silicon)—a magnetic dark gray powder—burns in air and chlorine.

Ziconium closely follows Ti but has been obtained in crystalline form, which looks like antimony but is harder—the oxide ZrO_2 is used for lime light.

POTASSIUM— K .

At. Wt. 39.—Val. I

History—

“Potash” from ashes known to ancients—not distinguished from “Soda” till middle of 18 century—metal K isolated by Davy in 1807.

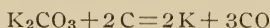
Occurrence—

The salts are found in rocks and all cultivated ground—In plants K occurs as oxalate and tartrate—In animals as chloride and phosphate—Sweat on sheeps wool is one source of K compounds—Largely obtained from “Argol” an acid, potassium tartrate ($\text{HKC}_4\text{H}_4\text{O}_6$) which mixed with coloring matter

deposits on side of wine casks—Another source is plant ashes which yield a crude K carbonate called POTASH—At Stassfort, large amounts are obtained from mineral carnallite (KCl).

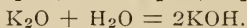
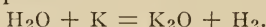
Preparation—

When Argol is heated in a closed retort it breaks up to calcium and potassium carbonate and carbon—on farther heating metallic K distills as a green vapor—essential reaction is



Properties—

A soft bluish white metal—green vapor—phosphoresces in dark—Of all the elements K has the greatest affinity for oxygen and chlorine at ordinary temperatures—it burns with a violet blue flame, but oxidizes in the air without taking fire—oxidizes on water, and decomposes it.



K and Sodium are used to reduce metals from their oxides.

Potassium Oxide K_2O —

The only important oxide of Potassium is K_2O , a gray solid formed by action of K on KOH—difficult to secure as it eagerly unites with H_2O forming KOH.

Potassium Hydrate KOH—

Formed when K acts on H_2O .

Generally prepared by treating a salt of K with hydrate of some metal which will form an insoluble salt with the radicle of the K salt.



Crude KOH obtained from wood ashes.

KOH is best example of an alkali and is strongest of all bases—it neutralizes acids to form salts—decomposes fats and oils to soaps hence destroys animal tissue—is soluble in $\frac{1}{2}$ its weight of water, forming a crystalline hydrate $\text{KOH} \cdot 2\text{H}_2\text{O}$.

KNO_3 (Salt peter).

Occurrence—

Spoken of under HNO_3 .

Preparation—

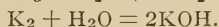
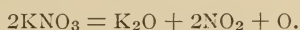
Usually by treating Chili saltpeter NaNO_3 with KCl .



NaCl being less soluble is removed by evaporation.

Properties—

Crystallizes in white, anhydrous, striated prisms, with cooling taste, and produces cold by its solution—large amounts act as poison—heated yields KNO_2 and oxygen, heated higher decomposes to KOH .



Melted in drops is “sal prunelle.”

KNO_3 oxidizes most substances save a few metals—with C it deflagrates.



Use—

Used in medicine and preparation of HNO_3 .

Chief use is in making gunpowder. Used instead of NaNO_3 because KNO_3 is not deliquescent. The theoretical equation in discharge of gunpowder is—



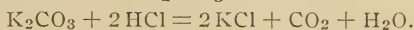
Theoretical proportions are—

KNO_3	S	C
74.8	11.8	13.4 parts.

POTASSIUM HALOIDS.

✧ Potassium chloride KCl —

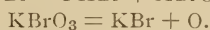
Called digestive salt, also sal febrifugium sylvii—found in sea-water and springs—occurs as sylvite and carnallite at Stassfurt—prepared from K_2CO_3 and HCl —



Crystallizes in colorless cubes—easily soluble in H_2O —solution absorbs heat—used largely in making alum, and K_2CO_3 —crude salt used as a fertilizer.

Potassium bromide KBr —

Colorless cubes—soluble in H_2O —made (1) directly from elements. (2) By action of Br on KOH —

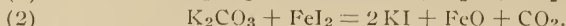
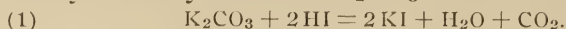


Usually (3) $\overset{\text{(Ferrous Bromide)}}{\text{FeBr}_2} + \overset{\text{(Ferrous Carbonate)}}{\text{K}_2\text{CO}_3} = 2\text{KBr} + \text{FeCO}_3.$

Used in medicine.

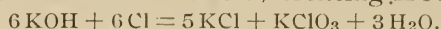
Potassium Iodide KI—

White crystalline solid, used in medicine and photography—usually made by action of K_2CO_3 on HI or FeI_2 —



Potassium Chlorate $KClO_3$ —

White crystalline solid made by treating KOH with Cl —



Substance treated under chlorine.

Potassium Perchlorate $KClO_4$ —

White solid formed by decomposition of $KClO_3$ —



Notable for its slight solubility.

Potassium Sulphide K_2S —

Potassium and sulphur form many compounds all soluble in water—chief is K_2S —prepared by fusing K_2SO_4 with C .



When fused is a red mass, but crystallizes from water solutions in colorless deliquescent prisms—"Hepar Sulphuris" or "liver of sulphur" is a mixture of the Polysulphides of K (K_2S , K_2S_4 , K_2S_5) and K_2SO_4 , obtained by fusing $K_2CO_3 + S$.

Potassium Sulphydrate KHS —

White, crystalline, solid, with alkaline reaction, prepared by action of H_2S on KOH .



Potassium Sulphate K_2SO_4 —

White, crystalline, solid—native as Kainite—prepared from



Soluble in H_2O —insoluble in absolute alcohol. Kainite is used as fertilizer.

Potassium Nitrite KNO_2 —

Formed by reduction of KNO_3 by heat.

Used as source of HNO_2 and in analysis.

Potassium Arsenite K_3AsO_3 —

Formed by action of K_2CO_3 on As_2O_3 .

A dilute solution used in medicine under name of "Fowler's solution."

Potassium Pyro=Antimoniate $K_2H_2SbO_7$ —

Made from KNO_3 and Sb.

Used in testing for Na.

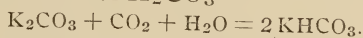
Potassium Carbonate K_2CO_3 —

Deliquescent salt—strongly alkaline—insoluble in alcohol.

Prepared by ignition of K salts of organic acids. Large quantities obtained from wood ashes—also from KCl as in the LeBlanc process for sodium carbonate, which see.

Potassium Hydrogen Carbonate $KHCO_3$ —

Called “Bicarbonate of potash” and prepared by passing CO_2 through a solution of K_2CO_3 —



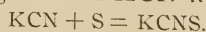
$KHCO_3$ is less soluble than K_2CO_3 and its solution gives a neutral reaction.

Potassium Cyanide KCN —

A poisonous salt—treated under Cyanogen.

Potassium Thiocyanate $KCNS$ —

Formed by treating molten KCN with S—



Used as a reagent for Iron (Fe).

SODIUM—NA—I.

History—

Metal discovered in 1807 by Davy.

Occurrence—

Not free—found in many minerals—occurs especially as NaCl in sea water, springs and as rock salt, and as $NaNO_3$ in Chili saltpeter.

Preparation—

Metal prepared as K, but easier, hence cheaper.

Properties—

Similar to potassium—silver-white metal with purple vapor—affinity for Cl, O, etc., only little less than that of K—decomposes H_2O with explosion*, and if water is warm burns with a yellow flame—forms with K the only liquid alloy not containing Hg—alloy has same appearance as Hg.

Hg + Na → NaHg
*The explosion is due to the globule of melted NaOH finally touching the H_2O , like water on a red-hot stove.

2 Na 2.170 + 1/2

SODIUM CARBONATE— $\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$.

Occurrence—

Found in nature in sea-plants just as K_2CO_3 is found in land plants.

Preparation—

Formerly obtained from ashes of sea plants—may be gotten by heating a solution of the primary carbonate, HNaCO_3 —



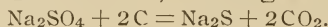
Commercially in two ways—

(1) LEBLANC's process from NaCl .

(a) NaCl to Na_2SO_4 by heating with H_2SO_4 —



(b) Na_2SO_4 to Na_2S by heating with C —



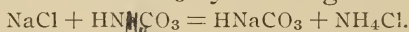
(c) Na_2S to Na_2CO_3 by heating with CaCO_3 —



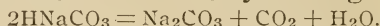
CaS is a waste product.

2) AMMONIA or SOLVAY process, which is more modern—

(a) NaCl to HNaCO_3 by treating with $\text{H}\overset{\text{Mono-am. carbonate.}}{\text{NH}_4}\text{CO}_3$.



(b) HNaCO_3 to Na_2CO_3 by heating.



The CO_2 is passed into ammonia forming again the acid ammonium carbonate $\text{H}\overset{\text{Mono-am. carbonate.}}{\text{NH}_4}\text{CO}_3$.

The NH_4Cl in (a) is heated with lime or magnesia and NH_3 set free as H_4NOH —into ammonia water thus formed, CO_2 from (b) is run forming $\text{H}\overset{\text{Mono-am. carbonate.}}{\text{NH}_4}\text{CO}_3$ again.

The Solvay process has no troublesome residue like the CaS of Le Blanc.

Properties—

Na_2CO_3 is an alkaline salt, efflorescent and very soluble in H_2O .

Use—

Essential in glass and soap making and a reagent in the laboratory.

Sodium Sulphate $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ —

“Glauber’s salt” occurs in nature especially in certain natural waters, as Carlsbad and Friedrichsbad springs—is a by-product in making HNO_3 .

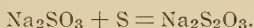


Crystallized in large colorless monoclinic crystals easily soluble in water and easily forming supersaturated solutions.

USED as a purgative in medicine and for producing artificial cold in the laboratory.

Sodium Thiosulphate $\text{Na}_2\text{S}_2\text{O}_3$ —

“Hypo” is a colorless crystalline solid easily soluble in H_2O —may be made by adding S to boiling solution of Na_2SO_3 .



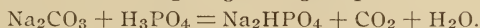
Chiefly used in photography and in bleaching as “anti-chlor.”

Sodium Nitrate NaNO_3 —

“Chili salt peter” is a deliquescent salt (see KNO_3) of value as the source of HNO_3 and KNO_3 , also used in coarser grades of gunpowder.

Di-sodium Phosphate $\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}$ —

White rhombic prisms—soluble in H_2O with slight alkaline reaction—formed from $\text{Na}_2\text{CO}_3 + \text{H}_3\text{PO}_4$ —



The crude H_3PO_4 containing CaSO_4 and H_2SO_4 can be used because Na_2SO_4 will not crystallize together with Na_2HPO_4 .

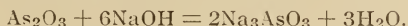
Acid Sodium Phosphate $\text{NaH}_2\text{PO}_4 \cdot \text{H}_2\text{O}$ —

White rhombic prisms easily soluble in H_2O , with strong acid reaction—chief cause of acid reaction in urine—made from Na_3PO_4 and H_3PO_4 —



Sodium Arsenite Na_3AsO_3 —

An uncrystallized substance formed from As_2O_3 and NaOH .



Sodium Arseniate Na_3AsO_4 —

Arseniate of soda forms in large crystals—prepared by ignition of As_2O_3 and NaOH with NaNO_3 .

Used in dissecting room as an antiseptic.

Sodium Tetraborate $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ —

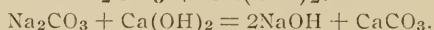
BORAX is formed as “tinkal” in India and California, but is mostly made artificially in Tuscany—heated, it loses water—melted, dissolves most of the metallic oxides, hence its use in soldering, bead and blowpipe work.

Sodium Oxides—

Na forms oxides like K, chief being Na_2O . Sodium oxide, which in preparation and properties resembles K_2O .

Sodium Hydrate NaOH —

Closely resembles KOH. Pure is prepared from $\text{Na} + \text{H}_2\text{O}$, commercially from $\text{Na}_2\text{CO}_3 + \text{Ca}(\text{OH})_2$.



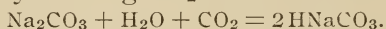
Purified by dissolving in alcohol.

Sodium Chloride NaCl —

“Common salt” is similar to KCl—Crystallizes in cubes without water—reaction of its solution is neutral—temperature makes but slight difference in its solubility—slightly soluble in alcohol—insoluble in HCl con.

Sodium Hydrogen Carbonate NaHCO_3 —

Saleratus or bicarbonate of soda is less soluble than Na_2CO_3 and is prepared by running CO_2 into a solution of carbonate—



Used in medicine and preparation of effervescing drinks.

Sodium Sulphide Na_2S —

Most important sulphide is Na_2S , sodium sulphide—analogous to K_2S —has alkaline reaction—formed by reduction of Na_2SO_4 with C—



Occasionally used in analysis.

Detection of Na salts—

Na salts detected by the yellow sodium flame or by precipitating as sodium antimoniate. All Na salts except NaNO_3 are efflorescent.

AMMONIUM— H_4N .

History—

The group H_4N has not been isolated, as it breaks up to $\text{H}_3\text{N} + \text{H}$ —after the discovery of Na and K attempts were made

at H_4N , but were unsuccessful—by decomposition of H_4N salt with Hg electrode, ammonium forms a voluminous unstable amalgam with Hg.

AMMONIUM CHLORIDE H_4NCl .

History—

“Sal ammoniac” was first made in Egypt by sublimation of camel’s dung.

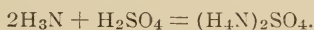
Occurrence—

Small amounts found native in volcanic regions

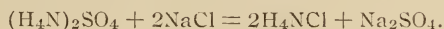
Preparation—

Formed from $\text{NH}_3 + \text{HCl} = (\text{H}_4\text{N})\text{Cl}$.

Obtained from the liquors of gas works (see NH_3)—the gas water is treated with H_2SO_4 .



$(\text{H}_4\text{N})_2\text{SO}_4$ is then sublimed with NaCl.



Properties—

Colorless, octohedral, crystals, highly soluble in water—causing cold by solution—volatile when heated by its dissolution into $\text{NH}_3 + \text{HCl}$. H_4NCl united easily with other salts to form double salts

The other haloid salts of H_4N are analogous, to H_4NCl .

USED in medicine and the arts.

H_4N SALTS.

Sulphate $(\text{H}_4\text{N})_2 \text{SO}_4$ —

Is amorphous with K_2SO_4 and highly soluble in water—its preparation from gas water given above.

Nitrate $(\text{H}_4\text{N}) \text{NO}_3$ —

Is amorphous with KNO_3 and deliquesces in air—decomposed by heat to N_2O and H_2O .



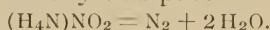
Highly soluble in water, forming a cooling mixture.

Nitrite $(\text{H}_4\text{N})\text{NO}_2$ —

White crystalline solid prepared by action of AgNO_2 on $(\text{H}_4\text{N})\text{Cl}$ —

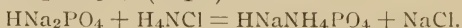


Its decomposition by heat yields pure N—



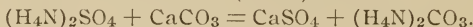
Ammonium=Sodium=Phosphate $\text{HNaNH}_4\text{PO}_4 \cdot 4\text{H}_2\text{O}$ —

“Microcosmic salt” is formed in stale urine—may be prepared from $\text{HNa}_2\text{PO}_4 + (\text{H}_4\text{N})\text{Cl}$ —



Carbonate—

White or transparent, hard mass giving off H_3N to the air—formed by action of CaCO_3 on $(\text{H}_4\text{N})_2\text{SO}_4$ —



Commercial salt contains also the bicarbonate, $\text{H}\text{NH}_4\text{CO}_3$, and the carbamate, $\text{NH}_2(\text{NH}_4)\text{CO}_2$, which are converted to $(\text{H}_4\text{N})_2\text{CO}_3$ by solution in $(\text{H}_4\text{N})\text{OH}$.

Thiocyanate $(\text{H}_4\text{N})\text{CNS}$ —

Prepared by action of CS_2 on H_3N .

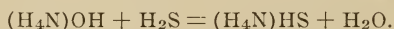


Decomposed by heat to thio-urea.

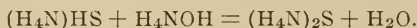


Sulphydrate $(\text{H}_4\text{N})\text{SH}$ —

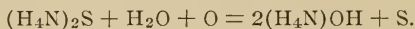
Prepared by running H_2S into ammonia water.



H_4NOH will now change it to $(\text{H}_4\text{N})_2\text{S}$.



The hydrosulphide dissolves S, forming the yellow ammonium sulphide, which contains the polysulphides and is valuable as a reagent— $(\text{H}_4\text{N})_2\text{S}$ decomposes gradually.



While any $(\text{H}_4\text{N})_2\text{S}$ remains, the S is dissolved, but finally precipitates.

LITHIUM.

History—

Discovered in 1017 by Arfoedsen in “PETALITE.”

Occurrence—

In small amounts quite widely distributed—chiefly as a compound silicate in lepidolite or “lithia mica,” found also in some mineral springs, and in ashes of many plants, notably tobacco and the beet.

Preparation—

Metal obtained by electrolyses of the chloride.

Properties—

Lightest of all metals (sp. gr. .59)—Oxidizes without melting in H_2O —Most Li salts are soluble in water and give a purple red color to the flame. Li characterized by the small solubility of Li_3PO_4 also of $LiCO_3$ (.75 in 100). Otherwise analogous to Na and K.

CÆSIUM—Cs—(sky-blue).

RUBIDIUM—Rb—(dark red).

At. Wt. 132.9. Val. I.

At. Wt. 85.4. Val. I.

Rare metals named from color of lines they give in spectro-scope, by means of which they were in 1860 discovered by Bunsen and Kirchhoff. Both are perfect analogues of K, and though rare are widely distributed. They frequently accompany K in mineral springs and plant ashes, and occur in larger quantities in lepidolite (Rb. 5%) and pollucite (Cs. 30%).

METALS OF THE ALKALIES.

<i>Symbol:</i>	Li.	H_4N .	Na.	K.
<i>At. Wt.:</i>	7	18	23	39.1
M_3SbO_4 :	soluble	insoluble	insoluble	soluble
$MHC_4H_4O_6$:	soluble	soluble	soluble	insoluble
M_2PtCl_6 :	soluble	insoluble	soluble	insoluble
MNO_3 :	deliquescent	deliquescent	deliquescent	permanent
M_2CO_3 :	difficultly sol.	soluble	efflorescent	deliquescent
M_3PO_4 :	difficultly sol.	insoluble	soluble	soluble
<i>Salts in general:</i>	permanent	volatile	efflorescent	deliquescent

CALCIUM—Ca.

At. Wt. 40. Val. II.

History—

In 1800 Davy obtained the metal from CaO .

Occurrence—

Widely distributed in large quantities—as CaO in most minerals—as CaCO₃ in limestone, chalk, marble—as CaSO₄ in gypsum, alabaster, etc.—as CaSO₄ and CaCO₃ in most natural waters—likewise contained as phosphate and carbonate in shells and bones of animals, and all ashes of plants.

Preparation—

Metal best obtained by electrolyses of CaCl₂.

Properties—

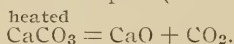
Yellow ductile metal—decomposes water—heated in the air Ca burns to CaO.

CALCIUM OXIDES.

Calcium forms two oxides, CaO and CaO₂.

CaO₂ is formed by heating CaCO₃ in stream of O.

The more important is CaO, commonly called “lime” or “quicklime.” Pure CaO is a white, infusible solid, best prepared by heating marble or Iceland spar (CaCO₃).



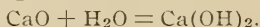
When exposed to moist air CaO again takes up water, forming Ca(OH)₂, which then unites with CO₂, forming the carbonate. By reason of its infusibility CaO gives with the compound blow-pipe the intense calcium light.

CALCIUM HYDRATE—Ca(OH)₂.

Preparation—

“Slacked lime” is formed by

(1) Adding H₂O to CaO.



Reaction is accompanied by considerable heat, sufficient to explode powder or char wood.

(2) Treating CaCl₂ with KOH.



Properties—

Soft white powder—attracts CO₂ of the air to form CaCO₃—at red heat the hydrate breaks up to CaO + H₂O. Ca(OH)₂ is slightly soluble in cold, and less in hot water, whence a sat-

urated cold solution becomes cloudy when heated. When CaO is added to water in excess, a portion of the hydrate goes into solution, the remainder is precipitated. The clear supernatant liquid is used as lime water, the bulky precipitate as "milk of lime."

LIME WATER has an alkaline reaction—unites easily with CO_2 —this is used as a test for CO_2 in the air—crystallized Ca(OH)_2 may be obtained from lime water—the solubility of Ca(OH)_2 is increased by presence of sugar and diminished by alkalis.

Slacked lime is used largely in ordinary mortar, a mixture of lime, water and sand, in which Ca(OH)_2 forms CaCO_3 from the CO_2 of air, and likewise forms with the sand a calcium silicate. HYDRAULIC mortar is produced by igniting limestone with aluminium silicate, the composition being chiefly Ca and Al silicates.

Ca(OH)_2 is cheapest of all bases and hence used wherever possible in preparing other hydrates.

CALCIUM CHLORIDE CaCl_2 .

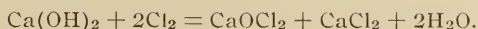
Made by treating CaCO_3 with HCl .



Some Ca(OH)_2 must be added to precipitate Fe or Mn present as impurities.—Chloride obtained by evaporation has formula $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ —is very soluble in H_2O and is used in freezing mixtures.—Anhydrous CaCl_2 is a white crystalline solid used to dry gases.

Bleaching Powder CaOCl_2 —

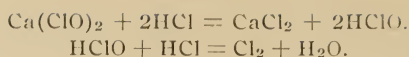
Treat slaked lime with Cl gas.



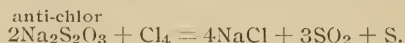
Formula of bleaching powder is undecided, but is approximately CaClOCl (CaOCl_2)—commercially called "chloride of lime," "chlorinated lime," "chemick."

Properties—

White porous powder—odor of chlorine— CO_2 of the air decomposes it, freeing HClO —by action of strong acids Cl is set free.



A weak acid as CO_2 sets free HClO —its strong bleaching and disinfection powers are due to these reactions—in bleaching, an alkaline bath, also an anti-chlor bath should follow.



CALCIUM SULPHATE $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$.

Occurrence—

Anhydrous forms occur as “gypsum” of which alabaster is one variety—natural waters contain CaSO_4 in solution.

Properties—

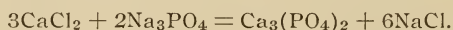
Not easily soluble in hot or cold water—presence of H_4N or Na salts or acids, markedly increases its solubility. Insoluble in $\text{C}_2\text{H}_6\text{O}$ (alcohol). Heated to 100° it loses most of its crystal water, forming PLASTER OF PARIS, an amorphous powder—this will again unite with water to form the crystalline variety—the “setting” is due to crystallization, which can be slowed by $\text{Na}_2\text{B}_4\text{O}_7$, and hastened by K_2SO_4 . If heated above 200° it becomes too dense to set.

CALCIUM PHOSPHATES.

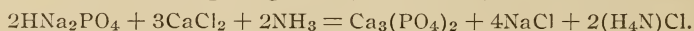
Tertiary or Neutral $\text{Ca}_3(\text{PO}_4)_2$ —

Most important natural occurrence is in bones (see P).

May be obtained from CaCl_2 and Na_3PO_4 .



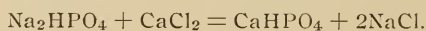
Or from the di-sodic phosphate (Na_2HPO_4) in alkaline solution.



When fresh is a gelatinous mass which dries to a white powder highly insoluble in water but easily converted to the soluble forms.

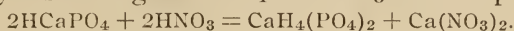
Secondary CaHPO_4 —

A white crystalline salt formed by treating Na_2HPO_4 with CaCl_2 .



Primary or Acid $\text{H}_4\text{Ca}(\text{PO}_4)_2$ —

White crystalline salt—soluble in H_2O with acid reaction—prepared by dissolving HCaPO_4 in HNO_3 and evaporating.



Largely used as a fertilizer.

Hypophosphite—

White crystalline salt—obtained by action of P on $\text{Ca}(\text{OH})_2$ —(see preparation of H_3P).

Calcium Nitrate $\text{Ca}(\text{NO}_3)_2$ —

A white deliquescent salt of little importance.

Calcium Carbonate CaCO_3 —

Widely distributed as limestone, chalk, marble, etc., purest in marble and Iceland spar—occurs likewise in forms of plant and animal life—insoluble in pure water, it dissolves somewhat in H_2CO_3 , hence found in all natural waters—Heated it readily yields CO_2 and CaO and is the chief source of Ca salts and CO_2 .

Calcium Sulphide—

Ca forms polysulphides of which CaS is the most important—Calcium sulphide is a yellowish mass, formed by heating CaSO_4 with C—chiefly notable from its phosphorescence, hence used in preparation of luminous paints.

REACTIONS OF CALCIUM SALTS.

NaOH and KOH: precipitate Ca salts as $\text{Ca}(\text{OH})_2$.

H_4NOH : gives no precipitate.

H_2CrO_4 : gives no precipitate (compare Ba).

H_2SO_4 : gives no precipitate (compare Ba and Sr).

$(\text{H}_2\text{N})_2\text{C}_2\text{O}_4$: precipitates in alkaline solutions.

P insoluble in $\text{H}(\text{C}_2\text{H}_3\text{O}_2)$ —soluble in mineral acids.

Flame: reaction is a red color when salt is moistened with HCl .

Blowpipe: Ca salts give a luminous residue.

STRONTIUM—Sr.

At. Wt. 87.5. Val. II.

The earth Strontia discovered in 1793—the metal by Davy in 1808 by electrolysis of the hydrate.

A rare metal found as the carbonate and sulphate.

Properties—

In general analogous to Ca.

Strontium Oxide SrO —

Prepared like CaO from the carbonate.

Strontium Hydrate $\text{Sr}(\text{OH})_2$ —

A strong base resembling $\text{Ca}(\text{OH})_2$ but more soluble in water. Forms salts like $\text{Ca}(\text{OH})_2$.

REACTIONS OF Sr SALTS.

NaOH and KOH: gives precipitate of $\text{Sr}(\text{OH})_2$ only in concentrated solution.
 H_2SO_4 : or soluble sulphates, precipitate SrSO_4
 Na_2CO_3 : precipitates SrCO_3 .
 Na_2HPO_4 : precipitates SrHPO_4 .

The only important salt is the nitrate, much used in fireworks.

BARIUM.

At. Wt. 137. Val. II.

History—

Metal obtained by Davy 1808.

Occurrence—

Chiefly as the carbonate (witherite) and the sulphate (barite).

Preparation—

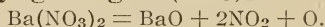
Best obtained by electrolysis of BaCl_2 . C will not reduce the metal from BaCO_3 , as it does with the alkali carbonates.

Properties—

Analogous to Sr and Ca, but a stronger base.

Barium Oxide BaO —

Prepared (1) by igniting $\text{Ba}(\text{NO}_3)_2$.



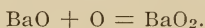
(2) Heating BaCO_3 with C.



Properties are similar to those of CaO .

Barium Peroxide BaO_2 —

Made by heating BaO in stream of O .



If strongly heated, BaO_2 yields its O .



This is one source of commercial O. BaO_2 is used also in preparing Ozone and Hydrogen peroxide.

Barium Hydrate $\text{Ba}(\text{OH})_2$ —



$\text{Ba}(\text{OH})_2$ crystallizes with 8 mols. of water, $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ —with water forms an aqueous solution called “Baryta water” which is a strong base and like $\text{Ca}(\text{OH})_2$ rapidly absorbs CO_2 from air—is considered a somewhat better reagent for carbonic acid.

REACTIONS OF Ba SALTS.

Barium salts similar to Ca salts but more poisonous.

NaOH :	precipitates $\text{Ba}(\text{OH})_2$ only in strong solutions.	} Differ from Ca and Sr.
H_4NOH :	gives no precipitate.	
H_2SO_4 :	precipitates the most dilute solutions.	
K_2CrO_4 :	insoluble	
Na_2CO_3 :	precipitates insoluble BaCO_3 .	

The sulphides of Ba, Sr, and Ca are all phosphorescent.

$\text{Ba}(\text{NO}_3)_2$ gives a green fire in pyrotechnics.

METALS OF THE ALKALINE EARTHS.

Symbol:	Ca.	Sr.	Ba.
At. Wt:	40	87.5	137
$M(\text{OH})_2$:	insoluble	medium	soluble
$M_2\text{CO}_3$:	insoluble	insoluble	insoluble
$M\text{Cl}_2$:	soluble	soluble	soluble
	soluble in $\text{C}_2\text{H}_6\text{O}$	soluble in $\text{C}_2\text{H}_6\text{O}$	insoluble in $\text{C}_2\text{H}_6\text{O}$
	deliquescent	permanent	permanent
$M(\text{NO}_3)_2$:	deliquescent	permanent	permanent
$M_2\text{SO}_4$:	1 : 400	1 : 6895	1 : 685,000
$M_2\text{CrO}_4$:	soluble	soluble	insoluble

MAGNESIUM—Mg.

At. Wt. 24. Val. II.

History—

In 1750 Black distinguished magnesia, MgO , from lime, CaO . In 1808 Davy showed that it was the oxide of a metal. Later Wöhler obtained the pure metal.

Occurrence—

Compounds less abundant than those of Ca but as widely distributed—contained in most soils and all plant ashes—in

large quantity in Hornblende and meerschaum (silicates) and dolomite (carbonate of Mg and Ca).

Preparation—

Metallic Mg prepared by electrolysis of the chloride heated to fusion, or heating the chloride with Na.

Properties—

Lustrous silver white metal—medium hard, ductile, and malleable. At ordinary temperature Mg gradually oxidizes on the surface—ignited burns with an intense light. Mg is easily attacked by dilute acids—slightly by the haloids—does not decompose H_2O , as $Mg(OH)_2$ is insoluble.

Uses—

Metallic Mg is made chiefly to burn, and comes next to lime in strength of light—used largely for “FLASHLIGHTS” in photography—forms many alloys, the one with Zn is often used in place of pure Mg as it burns with an equally bright light.

Magnesium Oxide MgO —“Magnesia”—

A white, very light powder, prepared by heating magnesia alba ($MgCO_3 \cdot Mg(OH)_2$). Commercial magnesia contains traces of SiO_2 , FeO and CO_2 .— MgO forms with H_2O an insoluble hydrate $Mg(OH)_2$.

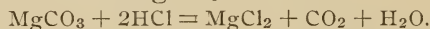
Magnesium Hydrate $Mg(OH)_2$ —

Formed as above, or by treating any Mg salt with $NaOH$.

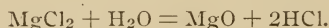
$Mg(OH)_2$ is a white amorphous powder, attracting CO_2 from the air—a strong base, almost insoluble in water, but soluble in $(H_4N)Cl$ —is used with $Fe_2(SO_4)_3$ as antidote for arsenic.

Magnesium Chloride $MgCl_2$ —

A deliquescent salt, occurring in sea water and mineral springs—prepared from $MgCO_3$ and HCl .



Heated, the deliquescent crystals yield water and break as below.



If H_4NCl should be present we would obtain anhydrous $MgCl_2$ —

The commercial salt occurs as a bi-product in making KCl and is much used in the arts.

Magnesium Sulphate $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ —

“Epsom salts” occur in sea-water and mineral springs, especially at Stassfurt. Pure MgSO_4 forms in rhombic efflorescent crystals, but usually contains some MgCl_2 . MgSO_4 is highly soluble in water with bitter saline taste—insoluble in alcohol.

Used—

In medicine—also to weight cotton cloth—a saturated solution with dextrine, will give a crystalized surface on glass.

Magnesium Carbonate MgCO_3 —

Occurs in compact masses as magnesite, also with CaCO_3 as dolomite.

The “magnesia alba” of medicine is a basic carbonate formed by treating an Mg salt with NaOH . Some CO_2 escapes and a white precipitate falls, which when dried at a low temperature has the formula $\text{Mg}(\text{OH})_2 \cdot 4\text{MgCO}_3 + 4\text{H}_2\text{O}$.

REACTIONS OF Mg SALTS.

$\text{Na}(\text{OH})$:	precipitates $\text{Mg}(\text{OH})_2$.
$(\text{H}_4\text{N})\text{OH}$:	“ “ incompletely; with $(\text{H}_4\text{N})\text{Cl}$ no precip.
Na_2CO_3 :	“ $\text{Mg}(\text{OH})_2 \cdot \text{MgCO}_3$.
$(\text{H}_4\text{N})_2\text{CO}_3$:	“ “ “ incompletely; never with $(\text{H}_4\text{N})\text{Cl}$
$(\text{H}_4\text{N})_2\text{C}_2\text{O}_4$:	no precipitate; compare Ca.
Na_2HPO_4 :	precipitate MgHPO_4 , which with H_4NCl changes to MgH_4NPO_4

BERYLLIUM.

At. Wt. 9. Val II.

A rare metal found native only as the oxide—common form is Beryl ($3\text{BeO}_2\text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$)—when green it is the emerald—Salts resemble those of Mg.

ZINC—Zn.

At. Wt. 64.9. Val. II.

History—

Brass or bronze was known to the ancients as evidenced in old coins—Zn was first recognized as a peculiar metal by Paracelsus who introduced the name, Zinc.

Occurrence—

Is comparatively rare—found native only in Australia—chief ores are “Smithsonite” (ZnCO_3) and zinc blende (ZnS)—Ores usually accompanied by Cadmium.

Extraction from ores—

By roasting in the air and then igniting the resulting oxide with C—



Zinc is quite volatile and hence the crucibles must be connected with a condenser—As in the case of S the Zn vapor condenses at first to a fine dust called Zinc dust, which contains Zinc, generally cadmium, and all the volatile impurities.—The commercial Zinc which forms after the dust, contains also many impurities especially Pb, Fe, C, sometimes S and Cd, and small amounts of As and Sb—it is purified by repeated distillation.

Detection—

Most characteristic compound is ZnS —before the blow-pipe it gives with $\text{Cu}(\text{NO}_3)_2$ a green color known as Rinman's green—the white oxide becoming yellow on ignition, is also characteristic.

Properties—

Blue white, crystalline metal—physical properties vary with temperature—brittle at ordinary temperature—at $100\text{--}150^\circ$ can be welded or drawn— 205° breaks under the hammer—melted and poured into cold water forms granulated Zn—not affected by air till strongly heated when it burns with strong blue flame to ZnO —pure Zn is slightly, commercial Zn is easily, attacked by acids and alkalis.

Used—

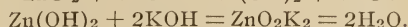
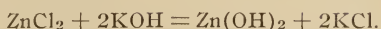
In galvanizing sheet-iron—forming alloys, especially brass—in galvanic batteries—with acids and alkaline hydrates as reducing agent.

Zinc Oxide ZnO —

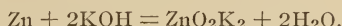
White powder, which becomes yellow with heat. Prepared as “Zinc white,” a pigment made by distilling metallic Zn—prepared for medical use by igniting the carbonate—for dental use by igniting the nitrate, the last two methods avoid As— ZnO occurs in nature as Zincite.

Zinc Hydrate $\text{Zn}(\text{OH})_2$ —

White powder—formed by treating Zn salt with KOH, and soluble in excess of alkali.



The same K zincate formed when metallic Zn is treated with KOH.



This reaction often used in alkaline reduction.

Zinc Chloride ZnCl_2 —

- 1) Heating Zn in current of Cl gas.
- 2) Dissolving Zn in HCl.

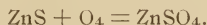


White deliquescent mass—fuses with heat—vaporizes at red heat—at high heat forms $\text{ZnO} + \text{HCl}$ (compare MgCl_2)—the ZnO unites with unchanged ZnCl_2 to form basic chloride—a concentrated solution of $\text{ZnCl}_2 + \text{ZnO}$ hardens with evolution of heat to ZnOHCl , which is used as a dental filling. ZnCl_2 is used as disinfectant, antiseptic, and wood preservative.

ZnBr_2 and ZnI_2 are analogous to ZnCl_2 .

Zinc Sulphate $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ —

“White vitriol” is found in Zn mines—prepared (a) by oxidation of blende (ZnS).



(b) by solution of Zn in H_2SO_4 .

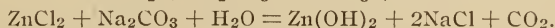
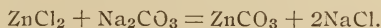


By ignition it forms the oxide ZnO .

Used in the arts and medicine, in battery fluids, etc.—used also to form colors by igniting with metallic salts—emetic.

Zinc Carbonate ZnCO_3 —

Formed in company with the hydrate, as basic Zn carbonate, when a Zn salt is treated with a soluble carbonate.



Notice tendency to form hydrates also in Mg.

Phosphates—

Most important are the basic phosphates—when metaphosphoric acid is mixed with ZnO it forms a cement, which is a basic metaphosphate.



Sulphide ZnS—

Zinc blende in nature is yellow—artificial is white—is readily formed from Zn dust and powdered sulphur. ZnS is soluble in mineral acids and alkalies—a little is formed by action of H₂S on ZnSO₄, but the acid freed soon dissolves it.

Zinc Salts—

Are colorless if the acid is colorless—normal and soluble salts redden litmus—Zn salts are poisonous, have metallic and a stringent taste and used as emetics.

CADMIUM—Cd.

At. Wt. 112. Val. II.

History—

Discovered in 1817 by Hermann and Stromeyer simultaneously.

Occurrence—

One of rarer metals—generally accompanies Zn ores—Zn blendes often contain from .2 to 3 percent Cd—the rare mineral “Greenockite” is CdS.

Preparation—

Cd is obtained in the first part of Zn distillate, i. e., zinc dust, and is purified by fractional distillation, as it is more volatile than Zn.

Properties—

A white, brilliant metal—malleable and tenacious, but the smallest trace of Zn makes it brittle.—Cd is not oxidized till at high temperature, when it burns to CdO.—Cd is soluble in H₂SO₄, HCl and best in HNO₃.—A pure Cd salt is precipitated completely by H₂S.—An amalgam of Zn + Cd is used as a dental filling.

Oxide CdO—

The only oxide of Cd is formed by burning the metal or igniting the nitrate or carbonate—the hydrates are formed by KOH acting on a Cd salt.

Sulphate CdSO_4 —

The sulphate $\text{Cd}(\text{SO}_4)_2$ is the most common salt of Cd and is anhydrous.

Nitrate $\text{Cd}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$ —

The nitrate $\text{Cd}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$ is very soluble and gives the oxide on ignition.

Chloride CdCl_2 —

The chloride CdCl_2 is similar to ZnCl_2 but efflorescent.

Cadmium Sulphide CdS —

The most important Cd salt occurs native as Greenockite—commonly made from H_2S and a Cd salt—not easily from the elements.— CdS is scarcely attacked by dilute acids, but decomposed by strong—forms a good yellow pigment—differs from As_2S_3 in volatility and solubility in alkalies.

CADMIUM SALTS.

In general resemble Zn salts.

NaOH and *KOH*: precipitate $\text{Cd}(\text{OH})_2$ insol. in excess (note difference from Zn).
(H₄N)OH: precipitates $\text{Cd}(\text{OH})_2$ soluble in excess.
Na₂CO₃: “ insoluble basic carbonate.
Na₂HPO₄: “ “ CdHPO_4 .
Na₂C₂O₄: “ “ CdC_2O_4 .
K₂CrO₄: does not precipitate dilute solutions (difference from Zn).
H₂S: precipitates yellow CdS insol. in acids (difference from Zn).
BaCO₃: precipitates CdCO_3 completely (difference from Zn).

COPPER—Cu.

At. Wt. 63.5. Val. II and I (-ic and -ous salts).

History—

Known to the ancients before iron and used for arms—name from Cyprus, where the Greeks and Romans got it.

Occurrence—

Native in large quantities—ores are generally the oxide and sulphide, especially chalcopyrite (CuFeS_2).

Extraction—

Two processes, the “English” and Mansfield, the latter used when ores are abundant but poor.—In English process the divided CuFeS_2 is first roasted in the air—this converts it partially to the oxide—then the ore is ignited with silica fluxes and carbon—the iron reduces to the oxide and is dissolved in the slag

—by repeating the process a blistered mass is obtained which contains much S—this is repeatedly roasted, melted and “poled” and Cu run into iron moulds. The Mansfield process is the same, save that a blast furnace is used and more silicious flux.

Properties—

Red crystalline metal—malleable—tenacious—weldable—conducts heat and electricity—not attacked by dilute acids save in presence of air—dissolves slightly in HCl, more in hot H_2SO_4 .



best in HNO_3 .



In dry air Cu is unaffected—in moist air, it forms CuO—with CO_2 is coated with CuCO_3 — H_2S blackens it, forming CuS.—Cu is blue green when melted, and expands on cooling—melted Cu always contains gas which escapes when it solidifies.—This is called “spitting” and prevents casting.

Pure Cu is used as wire, roofing, etc.—All casting must be made from alloys, of which brass, german silver and the various bronzes are examples.

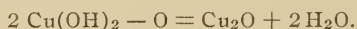
Copper forms four oxides, Cu_4O , Cu_2O , CuO, CuO_2 , of which Cu_4O is of least importance.

Copper Suboxide Cu_4O —

An olive green powder formed by reducing cupric hydrate.

Cuprous Oxide Cu_2O —

In nature as cuprite, and formed when cupric hydrate is reduced by grape sugar.



Cuprous hydrate $\text{Cu}_2(\text{OH})_2$ is a yellow precipitate which loses water on boiling to form the red Cu_2O . Cu_2O is used to give red color to glass.

CUPROUS SALTS

are not important and are easily converted to cupric form—they are generally colorless, red or yellow.

KOH :..... precipitates the yellow hydrate $\text{Cu}_2(\text{OH})_2$.

Na_2CO_3 :... precipitates the yellow carbonate Cu_2CO_3 .

$(\text{H}_4\text{N})\text{OH}$: dissolves salts colorless, but turns blue by oxidation.

Cupric Oxide CuO —

Found as the mineral cuprite—may be formed by oxidation of Cu or by heating the hydrate or nitrate—remarkable for its ease of reduction.—The hydrate $\text{Cu}(\text{OH})_2$ is a strong base which is changed to the oxide even when in water.

Copper Peroxide CuO_2 —

The peroxide is a yellow brown powder formed by treating $\text{Cu}(\text{OH})_2$ with H_2O_2 .

CUPRIC SALTS.

Generally blue or green when crystallized—colorless when anhydrous—color the flame blue or green—tartaric acid, sugar and many organic compounds prevent precipitation of $\text{Cu}(\text{OH})_2$.

NaOH : precipitates blue $\text{Cu}(\text{OH})_2$ (save as above).

H_4NOH : precipitates blue $\text{Cu}(\text{OH})_2$, this is soluble in excess of reagent and the solution thus formed dissolves cellulose.

Na_2CO_3 : precipitates basic CuCO_3 .

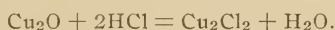
H_2S : precipitates black CuS .

$\text{K}_4\text{Fe}(\text{CN})_6$ precipitates red $\text{Cu}_2\text{Fe}(\text{CN})_6$.

Zn , Fe , Pb : precipitates metallic Cu from solutions.

Cuprous Chloride Cu_2Cl_2 —

The only cuprous salt of importance—formed from Cu_2O and HCl .



A white powder, insoluble in water,—absorbs carbon monoxide CO , hence is used in analysis.

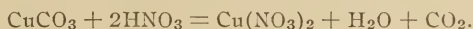
Cupric Chloride CuCl_2 —

Formed by dissolving the oxide or carbonate in HCl .



Cupric Nitrate $\text{Cu}(\text{NO}_3)_2$ —

Formed by action of HNO_3 on CuCO_3 .



Cupric Sulphate $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ —

Found native in mines, where CuS is oxidized—may be made by dissolving Cu in H_2SO_4 — CuSO_4 is soluble in three parts cold or one-half part hot water—

“Blue Vitriol” is much used in the arts in pigments, copper plating, galvanic batteries, etc.

Copper Arsenite CuHAsO_3 —

A greenish yellow precipitate, formed when cupric sulphate is added to potassium arsenite—commonly known as SCHEEL'S GREEN.

"PARIS GREEN" is also a copper arsenite, mixed with copper acetate. These colors are very poisonous and give off AsH_3 in the presence of organic matter—colors may be removed from fabrics by $(\text{H}_4\text{N})\text{OH}$.

Cupric Carbonate CuCO_3 —

The neutral salt CuCO_3 is not known—when Na_2CO_3 is added to Cu salt the basic carbonate is precipitated as either a blue precipitate $\text{Cu}(\text{OH})_2 + 2\text{CuCO}_3$ called "azurite" or a green $\text{Cu}(\text{OH})_2 \cdot \text{CuCO}_3$ called "malachite."

Cupric Ferrocyanide $\text{Cu}_2\text{FeC}_6\text{N}_6$ —

When acetic acid and potassium ferrocyanide are added to a Cu salt, a red brown precipitate of $\text{Cu}_2\text{FeC}_6\text{N}_6$ is formed, which constitutes the test for Cu.

Cupric Sulphide CuS —

A black compound found in nature, or precipitated by H_2S or $(\text{H}_4\text{N})_2\text{S}$ from solution of Cu salt. Insoluble in dilute acids—slowly changes in moist air to CuSO_4 .

ALLOYS OF COPPER.

The addition of Sn or Zn to Cu prevents "spitting" and forms an alloy harder than Cu, more durable, and of better color—likewise more sonorous.

Ancient bronze = Sn + Cu.—Modern = Cu + Sn + Zn.—German silver = Cu + Zn + Ni.—Brass = Cu + Zn, in different proportions and sometimes contains not more than .5 per cent of Pb, which facilitates working—a variety of brass alloys is produced by varying the amounts of Cu and Zn—the later increases hardness and fusibility, but decreases specific gravity and malleability. Aluminum bronze = Al + Cu in varying proportions.

In electro-plating, the article is immersed in a CuSO_4 bath—the negative pole of an electric battery is connected with the object— CuSO_4 decomposes and precipitates copper on the

object—in electro-typing, a mould of plaster of Paris is covered with graphite, and electro-plated as above.

SILVER—Ag.

At. Wt. 108. Val. I.

Occurrence—

Widely distributed—often native—the chief ore is the sulphide Ag_2S which occurs sometimes native, but generally with some other sulphides, especially galenite PbS . From argentiferous (silver bearing) lead ores, most commercial silver is obtained.

Extraction—

Three processes are in use: (1) (PATTINSON PROCESS.) Galena, a mixture of Pb and Ag sulphides, is strongly heated to melting—much of the lead separates out and some is united as PbAg alloy—the pure lead crystallizes first and is removed, leaving a readily fusible alloy of PbAg —by repetitions an alloy rich in Ag is obtained which is then cupelled. 2) (Parkes or Zinc process). The molten alloy is treated with Zn —the alloy of zinc and silver thus obtained is treated with superheated steam, which oxidizes the Zn , leaving Ag unchanged. 3) (Amalgamation process) Galena is roasted with NaCl , forming AgCl . H_2O and Fe reduce it to Ag , which is separated by amalgamating with Hg , and then distilling.

Properties—

The whitest metal—harder than gold and softer than Cu —very malleable and tenacious—at high temperature forms a light blue vapor—melted Ag absorbs O from the air and contracts on cooling—its alloys with Cu do not contract and polish better than Ag —silver is not oxydized at ordinary temperatures but O_3 changes it to the oxide Ag_2O — Ag is not affected by alkalis or KNO_3 —unites with Cl , Br and I at ordinary temperature and decomposes H_2S forming Ag_2S , which is soluble in KCN ,— Ag is soluble in HNO_3 , and H_2SO_4 conc., but AgNO_3 is insoluble in strong HNO_3 , hence dilute HNO_3 should be used in testing for chlorides.

Silver has three oxides, Ag_4O , Ag_2O and AgO , of which the first is of least importance.

Silver Suboxide Ag_4O —

A brown powder formed like Cu_2O and easily converted to Ag_2O and Ag_2 .

Silver Oxide Ag_2O —

The most important oxide of silver is a dark brown precipitate formed when KOH is added to AgNO_3 —any Ag hydrate formed is easily decomposed, leaving Ag_2O , which acts as if it were $\text{Ag}(\text{OH})$ —a strong base, slightly soluble in water and reduced by light, heat or H at 100° . Its soluble salts are poisonous and have metallic taste.

Peroxide AgO —

Black crystals with metallic lustre formed by passing ozone over Ag or Ag_2O .

Silver Chloride AgCl —

Found in nature as “Horn silver,” and formed when HCl is added to a soluble Ag salt— AgCl forms as a white, curdy precipitate insoluble in dilute acids, readily soluble in $(\text{H}_4\text{N})\text{OH}$, whence it crystallizes in large octohedrals—Fused AgCl sodifies to a horn-like mass, hence the name.

Silver Bromide AgBr —

Separates from silver salts on addition of HBr or a soluble bromide—a bright yellow precipitate similar to the chloride but less soluble in $(\text{H}_4\text{N})\text{OH}$.

Silver Iodide AgI —

Yellow, curdy precipitate formed from the action of HI or a soluble iodide, on silver salts in solution—insoluble in H_4NOH .

Photography—

Certain chemical compounds are sensitive to sunlight or other chemically active rays, (as Mg or P light)—Prominent among these are the Argentic haloids (AgX), which form violet changing to black argentous compounds (Ag_2X)—This is the basis of photography. AgCl is the most sensitive to light, but does not develop well; hence the “wet” plate is sensitized by a mixture of AgBr and AgI —The “dry” plate with AgBr . The exposed plate is treated with some “developer” like pyrogalllic acid, which reduces the argentous haloid (Ag_2Br) formed to metallic silver and dissolves free bromine—Metallic

silver unites with Argentic Bromide, forming argentous, which is again reduced by developer—Process continues till sufficient “molecular” silver is formed to give a distinct image on the plate—After sufficient development the plate is transferred to a “fixer” of sodium hyposulphite (“hypo”), which reacts with unchanged silver haloid, freeing Br and forming the soluble $\text{Na}_2\text{S}_2\text{O}_3 \cdot \text{Ag}_2\text{S}_2\text{O}_3$.

Silver Sulphate Ag_2SO_4 —

Obtained by dissolving Ag in hot H_2SO_4 —is used in refining of Ag—slightly soluble in H_2O .

SILVER NITRATE— AgNO_3 .

Preparation—

“Lunar caustic” is prepared by dissolving Ag in HNO_3 .

Properties—

A white crystalline solid, soluble in water, ether and alcohol—a brittle substance, but less so if accompanied by AgCl , which is generally present in the commercial article—when pure is unaffected by light but in presence of organic matter is reduced to metallic Ag.

Silver Fulminate $\text{Ag}_2(\text{CN})_2\text{O}_2$ —

Made by dissolving AgNO_3 in $\text{C}_2\text{H}_5\text{OH}$ and HNO_3 —more explosive than the Hg salt.

Cyanide AgCN —

AgCN is formed when HCN is added to solution of AgNO_3 —used in electro-plating.

Carbonate Ag_2CO_3 —

Grayish white precipitate, formed when Na_2CO_3 is added to a silver salt.

Chromate Ag_2CrO_4 —

Ag_2CrO_4 is precipitated by K_2CrO_4 from soluble Ag salt—a yellow precipitate, which is an indicator in analysis, where Ag salt is used.

Silver Sulphide Ag_2S —

Occurs native as Argentite, or precipitated by H_2S or $(\text{H}_4\text{N})_2\text{S}$ from Ag salts—a black precipitate used in silvering mirrors.

Alloys—

Silver is generally used as an alloy with copper—In coin 90% is Ag, in silverware about 75%, “sterling” silver 92%.

SILVER SALTS.

- NaOH* gives precipitate of Ag_2O .
(H₄N) OH..... gives precipitate of Ag_2O , soluble in excess; (in presence of free acid no precipitate).
Na₂CO₃..... gives precipitate of yellowish white Ag_2CO_3 .
Na₂HPO₄..... gives precipitate of yellowish Ag_2HPO_4 .
H₂S and (H₄N)₂S give precipitate of black Ag_2S .
HCl..... gives precipitate of white curdy AgCl [sol. in $(\text{H}_4\text{N})\text{OH}$, KCN , and $\text{Na}_2\text{S}_2\text{O}_3$ —insoluble in acids.]
KI..... gives precipitate AgI [insoluble in $(\text{H}_4\text{N})\text{OH}$, but soluble in $\text{Na}_2\text{S}_2\text{O}_3$.]
FeSO₄ gives precipitate of metallic Ag.
Tannic Acid..... gives precipitate of metallic Ag.
Zn and Fe..... gives precipitate of metallic Ag.
 All Ag salts are anhydrous.

MERCURY—Hg.

At. Wt. 200. Val. I and II.

History—

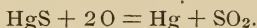
Mercury has been known since the earliest times.

Occurrence—

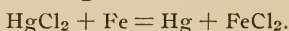
Not widely distributed nor in large quantities—comes mostly from Spain and California—native in small quantity—chief ore in Cinnabar HgS —Rare ores are “Amalgam” AgHg and and “Horn mercury” $2\text{Hg}_2\text{Cl}_2$.

Extraction—

Native Hg is filtered through leather—cinnabar is heated with O.



Commercial is sometimes pure, but often contains Pb, Sn, Bi and Cu dissolved—Pure Hg can be made from



Properties—

The only liquid metal—bluish white color, metallic lustre—at 40° crystallizes in octohedra—evaporates at ordinary temperatures—Hg resembles Au and Pt in its small affinity for O and large affinity for Cl—requires high temperature to form the

oxide, HgO , and at still greater heat yields it up again (compare BaO_2)—Ozone and chlorine attack it at normal temperature, but boiling HCl is harmless— Hg is easily divided by fats and chalk.

Used in manufacture of thermometers and barometers—tin amalgam for mirrors—extracting gold and silver from ores—useful also in electric connections.

Mercurous Oxide Hg_2O —

Mercury forms two oxides, Hg_2O and HgO — Hg_2O , corresponding to Ag_2O , is a heavy black powder formed from $\text{Hg}_2(\text{NO}_3)_2$ and KOH —easily decomposed by light to $\text{HgO} + \text{O}$, also by heat first to $\text{HgO} + \text{O}$, then to $\text{Hg} + \text{O}_2$.— HCl converts it to Hg_2Cl_2 ; Hg_2O is the basis of mercurous salts.

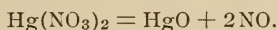
MERCUROUS SALTS.

KOH : precipitates black Hg_2O .
 $(\text{H}_4\text{N})\text{OH}$: precipitates black Hg_2O containing NH_3 .
 Na_2CO_3 : precipitates yellow unstable Hg_2CO_3 (which changes to the black oxide when heated).
 H_2S and $(\text{H}_4\text{N})_2\text{S}$: precipitates black Hg_2S .
 KI : precipitates green Hg_2I_2 .
 HCl : precipitates white Hg_2Cl_2 .
 SnCl_2 : precipitates white Hg_2Cl_2 (afterwards black Hg).
 Cu : precipitates metallic Hg .

Mercurous salts are mostly insoluble—have a metallic taste and act less violently than mercuric salts (Hg).

Mercuric Oxide HgO —

Is obtained as the red or yellow allotropic modification according to preparation—Red oxide by heating $\text{Hg}(\text{NO}_3)_2$.



Yellow oxide in the wet way from HgCl_2 and NaOH .



With heat HgO changes to the black oxide Hg_2O , but recovers its oxygen on cooling—The yellow and red differ slightly in chemical action; for example, the yellow is at once converted to the oxalate by $\text{H}_2\text{C}_2\text{O}_4$, but the red must be long heated—both are strong oxidizing agents and form the basis of mercuric salts.

REACTIONS OF MERCURIC SALTS.

KOH..... precipitates the yellow oxide HgO .

H₄NOH... gives a "white precipitate" in very dilute solutions.

KI..... gives a yellow precipitate of HgI , which turns red.

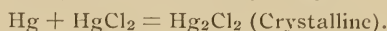
H₂S..... in large amounts gives black HgS .

H₂S..... in small amounts gives yellow precipitates $\text{HgX}_2 \cdot \text{HgS}$.*

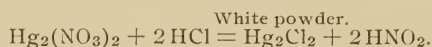
Cu..... is covered with a deposit of metallic Hg .

Mercurous Chloride Hg_2Cl_2 —

"Calomel" is prepared in (1) Dry way



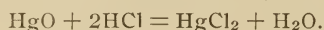
(2) Wet way



Hg_2Cl_2 sublimes without melting—somewhat less volatile than HgCl_2 —insoluble in H_2O , ether and alcohol—not easily attacked by acids, but boiled with HCl is converted to HgCl_2 .

Mercuric Chloride HgCl_2 —

"Corrosive sublimate" is formed (1) from HgO and HCl —



HgCl_2 is a highly volatile substance—soluble in cold water, better in hot—soluble also in alcohol or ether—aqueous solution reacts acid and in light decomposes to Hg_2Cl_2 —when boiled some of HgCl_2 is volatilized—mercuric chloride is precipitated by albumenoid substances, hence these are the best antidotes.

Mercurous Nitrate $\text{Hg}_2(\text{NO}_3)_2$ —

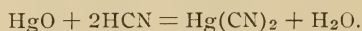
Made by action of dilute HNO_3 on excess of Hg —unless metallic Hg is present it gradually changes to $\text{Hg}(\text{NO}_3)_2$.

Mercuric Nitrate $\text{Hg}(\text{NO}_3)_2$ —

Solution may be obtained by dissolving Hg , or HgO in excess of strong HNO_3 ; acid must be in excess or basic salt will separate—the basic salt is reconverted by boiling in H_2O .

Mercuric Cyanide $\text{Hg}(\text{CN})_2$ —

Chief source of $(\text{CN})_2$ gas and is prepared from HgO and HCN .



*X stands for any non-metal.

Mercuric Fulminate $\text{HgC}_2\text{N}_2\text{O}_2$ —

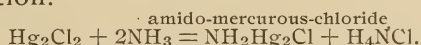
Analogous to the Ag salt but less explosive—used in percussion caps.

Mercuric Sulphide HgS —

Occurs in Cinnabar and Vermillion—Is produced as black amorphous mass, by action of H_2S on a mercuric salt—If this black sulphide is heated it sublimes as a red crystalline mass.

SUBSTITUTED H_4N COMPOUNDS.

When H_4NOH is poured over Hg_2Cl_2 it blackens with the following reaction.



The new compound is regarded as ammoniumchloride in which two hydrogens are replaced by Hg_2 . This one of several substitution products formed by the entrance of Hg into ammonium compounds.

METALS OF THE Cu GROUP.

	Cu	Ag	Hg
SPECIFIC GRAVITY.....	8.95.....	10.57.....	13.59
MELTING POINT.....	1054°	954°	39°
AT. WT.....	63.4.....	107.6.....	199.8

Ag resembles somewhat the alkali metals—group chlorides with symbols MCl_2 are soluble in H_2O .

Chlorides M_2Cl_2 are insoluble.

Sulphides are black, and insoluble in acids or alkalies but CuS is slightly soluble in $(\text{H}_4\text{N})_2\text{S}$.

LEAD—Pb.

At. Wt. 208. Val. II.

History—

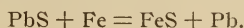
One of earliest known metals.

Occurrence—

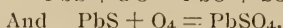
Occasionally native—chief ore is Galena PbS , a very common and widely distributed mineral.

Extraction—

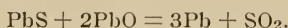
(1) By heating the sulphide with Fe.



(2) By roasting until the sulphide is partially converted to PbO and PbSO₄.



By ignition, the lead now separates.



Pure lead is obtained by heating PbC₂O₄ with C.

Properties—

Lustrous, blue white metal—at 1,600° gives off poisonous vapors—all its salts are poisonous—at red heat is easily oxidized—at normal temperature, a thin coat of the oxide forms, which protects the body of metal.—Pb is dissolved by HNO₃ but not by HCl, nor dilute H₂SO₄ because of insoluble salts formed—From its nitrate, Pb(NO₃)₂, lead is precipitated in metallic form, by Zn, forming a LEAD TREE. When Pb is in contact with air and water the somewhat soluble hydrate Pb₂O(OH)₂ is formed. This is especially dangerous where pure H₂O or rain water STANDS IN lead cisterns and pipes. With hard water the action is very slight.

Salts and organic matter modify, CO₂ and CaH₂(CO₃)₂ and sulphates diminish, chlorides and nitrates and especially nitrogenous organic matter increase, corrosion. Lead in contact with wood is rapidly corroded.

Lead Suboxide Pb₂O₃—

Lead forms four oxides, Pb₂O, PbO, Pb₂O₃ and PbO₂.

Pb₂O is the gray coating on metallic Pb—it decomposes to Pb + PbO.

Lead Oxide PbO—

Found occasionally in nature—artificial is of various colors according to preparation, *i. e.*, red, yellow and white, due probably to crystallization.

By ignition of Pb we get the yellow which becomes red when rubbed—when strongly heated becomes yellowish red, forming “litharge”.—PbO is a strong base resembling BaO and SrO—it absorbs CO₂ from the air, forming PbCO₃.

Lead oxides are much used in lead plaster, glazing earthen ware, making glass, etc.

Lead Trioxide and Dioxide Pb_2O_3 , PbO_2 —

Pb_2O_3 is of no importance. PbO_2 is formed on positive pole in the electrolysis of a lead salt—also by action of Cl on PbCO_3 —a dark brown powder, which conducts itself somewhat like MnO_2 , as for example, when heated with HCl, chlorine gas is formed.

Minium Pb_3O_4 —

“Red lead” is a mixture of PbO and Pb_2O_3 —important as a pigment, and made by oxidizing Pb.

Lead Hydrate $\text{Pb}_2\text{O}(\text{OH})_2$ —

When a Pb Salt is treated with KOH it does not form $\text{Pb}(\text{OH})_2$, but a basic hydrate of approximately $\text{Pb}_2\text{O}(\text{OH})_2$ —In presence of strong bases, $\text{Pb}_2\text{O}(\text{OH})_2$ acts as an acid, hence is soluble in KOH.

Lead Chloride PbCl_2 —

A WHITE PRECIPITATE formed when HCl, or a soluble chloride, is added to a cold solution of Pb salt. Soluble in hot water, but is largely deposited in needle shaped crystals when cool—melted PbCl_2 solidifies to a horn like mass.

Bromide PbBr_2 —Iodide PbI_2 —

PbBr_2 is very similar to the chloride—chief characteristic of PbI_2 is that it crystallizes from its hot water solution, in yellow lustrous laminae.

Lead Sulphate $\text{Pb}(\text{SO}_4)$ —

Found in nature, or prepared by adding a soluble sulphate to a lead salt—almost insoluble in water and dilute H_2SO_4 —soluble in HNO_3 and $(\text{H}_4\text{N}) \text{C}_2\text{H}_3\text{O}_2$.

Nitrate $\text{Pb}(\text{NO}_3)_2$ —

The most common soluble salt of PbO is formed by dissolving lead or its oxide in HNO_3 .

Lead Chromate PbCrO_4 —

A yellow precipitate, formed when a soluble chromate reacts on a soluble Pb salt—it is the test for either—it is used as a pigment and is highly poisonous from both of its ingredients.

Sulphide PbS —

“Galena” is the chief ore of lead—occurs in cubical crystals with metallic lustre—is precipitated by H_2S from Pb salt.

White Lead—

Is one of the oldest pigments—A basic lead carbonate of variable composition, made by passing CO_2 through solution of $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ (French method), or exposing rolls of Pb to atmosphere of $\text{H}(\text{C}_2\text{H}_3\text{O}_2)$ and CO_2 —often adulterated with BaSO_4 and CaSO_4 .

LEAD SALTS.

NaOH :.....	precipitates $\text{Pb}_2\text{O}(\text{OH})_2$, soluble in excess.
H_4NOH :.....	“ $\text{Pb}_2\text{O}(\text{OH})_2$, insoluble in excess.
Na_2CO_3 :.....	“ white PbCO_3 , soluble in excess.
H_2S :.....	“ black PbS .
KI :.....	“ yellow PbI_2 , soluble in hot H_2O .
K_2CrO_4 :.....	“ yellow PbCrO_4 .
HCl :.....	“ PbCl_2 , in strong solutions.
H_2SO_4 :.....	“ PbSO_4 , insoluble in H_2O or dilute H_2SO_4 —soluble in H_2SO_4 concentrated.
$\text{H}_4\text{NC}_2\text{H}_3\text{O}_2$:	decomposes PbSO_4 .
<i>Fe and Zn</i> :...	separate crystalline Pb from solutions.
<i>Blowpipe</i> :....	gives a yellow coating and malleable lead.

THALLIUM—Te.

At. Wt. 204. Val. I or III.

A spectrum metal found by Crookes in the mud from Swedish H_2SO_4 chambers. Occurs in minute quantities in combination with other metals—chemical character is partly like Na, partly like Pb. Oxidizes rapidly in the air, forming Te_2O and Te_2O_3 —most salts are soluble.

ALUMINIUM—Al.

At. Wt. 274. Val. III. or IV.

History—

In 1817 the metal was first obtained by Wöhler from Al_2Cl_6 and K.

Occurrence—

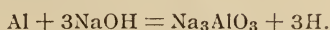
Widely distributed in large quantities as the oxide, in ruby, sapphire, corundum and emery—as silicate, in clay, mica and most crystalline rocks.

Preparation—

- (1) By fusing the chloride with metallic Na.
- (2) Electrolysis of the chloride, Al_2Cl_6 .

Properties—

Silver white, highly malleable, ductile and sonorous—a light, tenacious metal—in bulk, Al is stable in the air—easily soluble in HCl, or boiling H_2SO_4 , not in HNO_3 . NaOH dissolves it forming an aluminate.



Unites directly with S group, P, As, Si, and Haloids. In foil, Al burns in the air, and decomposes H_2O . At 100° Al alloys easily with some metals, especially with Cu in aluminium bronze.

Metallic Al is used in scientific instruments, watch-springs, ornaments, etc.—Aluminium bronze, in watches, spoons, etc.—Aluminium silicate or clay, in pottery, bricks, etc.—Aluminium acetate, as a mordant in dyeing cotton cloth.

AL COMPOUNDS.

Aluminium Oxide Al_2O_3 “Alumina”—

Found in nature as ruby, sapphire, etc.—prepared by heating the hydrate $\text{Al}_2(\text{OH})_6$ or igniting an alum—The oxide is infusible, save in oxyhydrogen flame—if ignited, is insoluble in H_2O and acids—before ignition, is soluble in acids, and KOH.

Aluminium Hydrate $\text{Al}_2(\text{OH})_6$ —

A voluminous, white precipitate, formed when $(\text{H}_4\text{N})\text{OH}$ is added to solution of an Al salt—If any vegetable coloring matter should be present, it would be precipitated with the hydrate forming an insoluble compound, technically known as a “LAKE”— $\text{Al}_2(\text{OH})_6$ freshly formed, dissolves in excess of KOH or in acids. After long standing its solubility diminishes—heated, it forms Al_2O_3 .

Aluminium Chloride Al_2Cl_6 —

Formed when Cl gas is passed over hot Al—the hydrated salt is formed by dissolving $\text{Al}_2(\text{OH})_6$ in HCl— Al_2Cl_6 is a colorless deliquescent salt—forms double salts with other metals—aqueous solution is a disinfectant.

Aluminium Sulphate $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$ —

A white crystalline solid, obtained from solution of $\text{Al}_2(\text{OH})_6$ in H_2SO_4 , or industrially, from clay and H_2SO_4 —easily soluble

in water—becomes anhydrous with heat—at high temperature decomposes to $\text{Al}_2\text{O}_3 + \text{SO}_4$.

Alums—

Aluminium sulphate combines with the alkali sulphates to form “alums.” The chief alum is potassium aluminium sulphate $\text{K}_2\text{Al}_3(\text{SO}_4)_4 + 24\text{H}_2\text{O}$, “common alum.” Alum is a general term, applied to a series of double salts of similar composition and crystalline form. Iron, chromium and manganese form similar derivatives by replacing the Al_2 —a further series is formed, by replacing K_2 with Na_2 or $(\text{H}_4\text{N})_2$, thus we have iron alum, $\text{Fe}_2\text{K}_2(\text{SO}_4)_4 + 24\text{H}_2\text{O}$, chrome alum $\text{Cr}_2\text{K}_2(\text{SO}_4)_4 + 24\text{H}_2\text{O}$, and others of similar formula.

ALUMINIUM SALTS.

The haloids, sulphates, nitrates and acetates are soluble.

NaOH or KOH: gives $\text{Al}_2(\text{OH})_6$ soluble in excess.

H₄NOH: gives $\text{Al}_2(\text{OH})_6$ insoluble in excess.

Na₂CO₃: “ “ “ “ “

(H₄N)₂S “ “ “ “ “

AL SILICATES.

Occurrence—

Granite contains both Al silicate and alkaline silicates. When it disintegrates, the alkaline silicates are washed away, leaving insoluble aluminum silicate, clay—purest form of clay in Kaolin—Topaz, Beryl and Lapis Lazuli, are important silicates, especially the latter, of which the powder was the old source of Ultramarine.

Ultramarine—

Is now made by heating a mixture of clay, Na_2CO_3 , S and C.—this gives a colorless compound, turning green—by gently heating with sulphur, it forms the blue variety—variation in heat and ingredients gives a variety of colors.

Porcelain—

Is essentially an Al glass made from kaolin—true porcelain is translucent—slightly attacked by reagents.

Pottery—

Is an impure porcelain made from clay—its red color due to FeO .

GLASS.

Properties—

A mixture of silicates of the alkalis and other metals—a thin liquid at high temperatures, becoming viscous as it cools. Amorphous, transparent solid, little affected by acids or water. No single silicate has all these properties. The alkali silicates are amorphous and transparent but soluble. Others are insoluble, but crystalline. A combination has the required properties. There are four classes of glass.

I—Bohemian or Hard Glass—

K and Ca silicate, which melts at high temperature and resists reagents—Used in laboratories.

II—French Glass—“Crown Glass”—

Na and Ca silicate of blue-green color, harder than Bohemian but easier melted. Another name is “crown glass.”

III—Bottle Glass—

Impure glass, colored by Fe_2O_3 —a silicate of the alkalis, Ca, Mg and Al.

IV—Lead or Crystal Glass—

K and Pb silicates—Pb increases refraction, hence used as “strass” glass to imitate jewels. Used in optics as “flint” glass.

Colored Glass—

Contains Cu_2O (red), or Au(ruby), CuO (green), Co(blue), Mn(violet), etc.

GALLIUM GA.

At. Wt. 70. Val. III or IV.

INDIUM IN.

At. Wt. 113.4. Val. III or IV.

Two metals discovered through the spectroscope, and closely analogous to aluminium—the aluminium group, Al, Ga, In, are true metals—their oxide M_2O_3 , is a weak base, acting as an acid in presence of strong alkalis—the metals do not decompose water—are easily dissolved by haloid acids with evolution of H. Their valence is III. or IV.



Alums are easily formed from their sulphates.

MANGANESE—Mn.

At. Wt. 55. Val. II. and IV.

Occurrence—

Never free—chief one is Pyrolusite, (MnO_2)

Extraction—

By igniting the oxide with C.

Properties—

Hard, gray, metal—fuses with difficulty—not stable in the air—metal has no technical use—forms an alloy with Fe which is used in Bessemer process.

OXIDES OF MANGANESE.

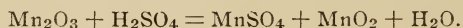
MnO , Mn_2O_3 , Mn_3O_4 , MnO_2 , Mn_2O_7 .

Manganous Oxide MnO —

Formed by reduction of higher oxides or by ignition of the oxalate—reduced by C to the metal— MnO is the basis of common manganese salts—its hydrate $\text{Mn}(\text{OH})_2$ is a white precipitate easily oxidized—like $\text{Mg}(\text{OH})_2$ it is not precipitated in H_4NCl solution.

Manganic Oxide Mn_2O_3 —

A black powder—made by igniting the other oxides in oxygen—its hydroxide, $\text{Mn}(\text{OH})_3$ is prepared from $\text{Mn}(\text{OH})_2$. Treated with dilute acids, the oxide or hydroxide gives manganous salts.



Mn_2O_3 occurs as mineral “braunite.”

Mangano-Manganic Oxide Mn_3O_4 or $\text{MnO} \cdot \text{Mn}_2\text{O}_3$ —

Found as a red-brown powder, by ignition of the carbonate—acts like a mixture of $\text{MnO} + \text{Mn}_2\text{O}_3$ —is isomorphous with magnetite, Fe_3O_4 .

M_3O_4 occurs as mineral “hausmanite.”

Manganese Dioxide MnO_2 —

Occurs as “pyrolusite”—made by igniting manganous nitrate $\text{Mn}(\text{NO}_3)_2$ — MnO_2 is a black powder, which, heated, yields O—treated with HCl , gives off chlorine—its hydroxide is $\text{Mn}(\text{OH})_2$, from which the unstable manganites are derived.

OTHER MN COMPOUNDS.

Manganous Ammonium Phosphate MnH_4NPO_4 —

Like the Mg salt is formed when H_4NOH and alkaline phosphate are added to a Mn salt.

Manganous Carbonate MnCO_3 —

Precipitated by alkaline carbonates from manganous salts—a white powder which easily oxidizes turning brown—like CaCO_3 , is soluble in natural waters which contain CO_2 .

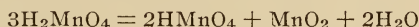
Manganous Sulphide MnS —

Precipitated from manganous solution by the alkaline sulphides—a flesh colored precipitate turning brown, then green. MnS_2 corresponding to MnO_2 is likewise found in nature as “hauerite.”

MN ACIDS.

Manganic Acid H_2MnO_4 —

The oxide MnO_3 is not known—The hydrated oxide, manganic acid H_2MnO_4 breaks up in solution to permanganic acid and MnO_2 .



Manganic acid is dibasic, corresponding to H_2SO_4 or $\text{H}_2\text{C}_2\text{O}_4$ and the salts of these three acids are isomorphous.

Potassium Manganate K_2MnO_4 —

The K salt is formed from $\text{MnO}_2 + \text{KOH}$.



Manganates are of green color and easily decomposed by acids to permanganates.

Permanganic Acid HMnO_4 —

Known both in salts and free state—the free acid is made from its Ba salt. HMnO_4 is a powerful oxidizing agent—decomposes above 40° —salts are purple and generally made by decomposition of manganates with acids.

Potassium Permanganate KMnO_4 ($\text{K}_2\text{Mn}_2\text{O}_8$)—

Chief salt of permanganic acid, is formed when CO_2 acts on K manganate, till the green color changes to red.



KMnO_4 crystallizes in dark red prisms—isomorphous with KClO_4 .

A powerful oxidizing agent—in oxidizing it is reduced to the colorless manganate, hence the persistence of the permanganate color tells when an oxidation is complete—Na-permanganate is used commercially as a disinfectant under name of CONDY'S FLUID.

OTHER MN COMPOUNDS.

Manganous Chloride MnCl_2 —

Red deliquescent crystals—formed when Cl is made from MnO_2 and HCl .



Manganous Sulphate MnSO_4 —

Bright red or pink crystals, formed from H_2SO_4 and MnO .



If higher oxides are used, oxygen will separate.

MANGANOUS SALTS.

Color—reddish when crystalline—otherwise colorless.

NaOH and KOH give white $\text{Mn}(\text{OH})_2$ becoming brown.

Na_2CO_3 gives “ “ “ “

BaCO_3 precipitates “ only from MnSO_4 .

$(\text{H}_4\text{N})_2\text{S}$ “ MnS , flesh-colored, turning green.

H_2S “ MnS but slightly, in neutral solution.

KNO_3 and

Na_2CO_3 heated with -ous salts give K_2MnO_4 (characteristic).

$\text{Na}_2\text{B}_4\text{O}_7$ “ “ “ amythyst color.

IRON—Fe.

At. Wt. 56. Val II. and III.

History—

The use of iron dates back to the early legends, when the source was probably meteoric. Later it was obtained from ores, as the great slag fields of India give witness.

Occurrence—

The most widely distributed metal—seldom found native save in meteors, where it generally contains some nickel—occurs in the blood and mineral waters—chief ores are magnetite Fe_3O_4 , hematite Fe_2O_3 , limonite $\text{Fe}_2\text{O}_3 + \text{H}_2\text{O}$, and siderite FeCO_3 .

IRON ORES.

Magnetite Fe_3O_4 —(Fe_2O_3 , FeO)—

Is the richest and best of iron ores—occurs in crystalline and amorphous forms—is reduced with difficulty, but yields excellent iron—magnetite may have Fe_2O_3 replaced by Mn_2O_3 , and FeO by ZnO —it is then an ore of zinc called franklinite—or may have its Fe_2O_3 replaced by Cr_2O_3 , and is then chromite.

Hematite Fe_2O_3 —

One of the most common ores of iron—occurs also in a crystalline modification known as specular iron ore.

Limonite—

Or bog iron is a mixture of Fe_2O_3 and $\text{Fe}_2(\text{OH})_6$ —yellow ochre is the clayey variety.

Siderite—

Known also as spathic iron ore, and when mixed with clay as clay iron stone, is the chief English ore.

Extraction—

Ores are pulverized and roasted, to drive out water, carbonic acid, or sulphur, and to convert oxides to ferric oxide Fe_2O_3 , which easily reduces. Ores are then heated with C and fluxes in the blast furnace. The furnace is charged with alternate layers of fuel (charcoal, coke or anthracite), ore and flux—a powerful blast of hot air oxidizes the C to CO_2 —this is reduced by hot C farther up the pile to CO—CO passing through the hot ore reduces Fe_2O_3 to a spongy mass of metallic Fe, which, mixed with flux and earthy impurities, settles to hotter part of furnace. Here the iron forms a fusible compound with C, and drops to bottom of furnace to be drawn off—the flux and impurities melt to a liquid “slag” which floats above the molten iron.

VARIETIES OF IRON.

Pig-Iron or Cast-Iron

is the crude iron as obtained from the furnace and may contain carbon, phosphorus, sulphur, silicon, manganese, etc. If cooled rapidly when taken from the furnace most of its carbon remains in combination and it is then known as WHITE CAST IRON—this leaves no residue when dissolved in acid, as all the C unites

with the H set free. If the iron is cooled slowly the C separates as graphite and is so left when iron is dissolved—this variety is GRAY CAST IRON. If the ore contains much manganese, this is reduced at same time and such manganese-containing iron, may combine with a greater quantity of carbon, and is known as SPIEGEL IRON. The presence of C, Si and P renders cast iron brittle, hence they must be removed before it can be welded—these substances are oxidized in the puddling furnace, which yields WROUGHT IRON containing less than .6 per cent of C.

Steel—

Is intermediate between cast and wrought iron—may be made by adding C directly to wrought iron as in the CEMENTATION process, which produces very good steel, or in BESSEMER process, by decarbonizing cast iron to form wrought iron, then adding enough cast iron to produce the desired percentage of C for steel.

Properties—

Commercial iron is never pure.—The several varieties differ chiefly by reason of impurities, especially C as given above—purest form is wrought iron, especially in piano wire which contains but three per cent. of impurities.

PURE FE may be obtained by igniting the oxide in a current of H—it is then a silver white metal—softer and more malleable than wrought iron, but less tenacious.

Iron oxidizes easily in moist air or under water to Fe_3O_4 —finely divided it decomposes H_2O at 100° —dilute acids dissolve Fe with evolution of H—concentrated HNO_3 does not affect it, due probably to thin coat of the oxide— CO_2 in water forms $FeCO_3$ —Fe unites directly with haloids with S, C, Si, and the metals.

OXIDES OF Fe.

Iron forms four oxides FeO , Fe_2O_3 , Fe_3O_4 , and the hypothetical FeO_3 . FeO and Fe_2O_3 are basic and form salts— FeO , a strong base, corresponds to CaO — Fe_2O_3 , a weak base, corresponds to Al_2O_3 — FeO_3 , like MnO_3 , exists only in the K salt.

Ferrous Oxide FeO —

Does not occur free but always as a salt, in nature—difficult to make, as it absorbs O—may be formed by reducing Fe_2O_3 with H—a black powder.

Ferrous Hydrate $\text{Fe}(\text{OH})_2$ —

The base of ferrous (Fe) salts—a white precipitate thrown out of ferrous salts by action of the alkalis in a current of H. On exposure to air it oxidizes to green, then red-brown color—insoluble in water—a strong base.

FERROUS SALTS—Fe.

White when anhydrous—blue-green when hydrated.

KOH and NaOH form white $\text{Fe}(\text{OH})_2$, easily oxidized.

K_2CO_3 precipitates white FeCO_3 , easily oxidized.

H_2S “ black FeS , only in dilute neutral solutions.

$(\text{H}_4\text{N})_2\text{S}$ “ “ FeS , soluble in acids.

$\text{K}_4\text{FeC}_6\text{N}_6$ “ white $\text{Fe}_2\text{FeC}_6\text{N}_6$, turns blue by oxidation.

$\text{K}_6\text{Fe}_2\text{C}_{12}\text{N}_{12}$ “ blue $\text{Fe}_3\text{Fe}_2\text{C}_{12}\text{N}_{12}$ “Turnbulls blue.”

KCNS gives no color when no Fe_2 salts are present.

Na_2HPO_4 precipitates white $\text{Fe}_3(\text{PO}_4)_2$, turns blue by oxidation.

Tannic acid gives no color when no Fe_2 salts are present.

$\text{Na}_2\text{B}_4\text{O}_7$ gives a colorless or yellow bead.

Ferrous salts absorb NO in the cold, turning black, and are used as a reagent for HNO_3 .

Ferric Oxide Fe_2O_3 —

Occurs in nature as hematite—prepared by ignition of ferrous sulphate—



Thus prepared Fe_2O_3 is a dark red powder, used as a paint under name “Colcothar”— Fe_2O_3 is a very hygroscopic substance and like MnO_2 aids in preparation of O from KClO_3 .

Ferric Hydrate $\text{Fe}_2(\text{OH})_6$ —

The basis of ferric salts—red brown voluminous precipitate, made by treating a ferric salt with an alkali hydrate—red heat renders it insoluble in acids..

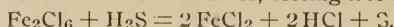
FERRIC SALTS.

Color—

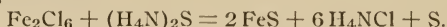
If anhydrous the neutral (Fe_2) salts are without color—when crystallized they are yellow or reddish yellow, save the nitrate and fluoride which are colorless, and the acetate and thiocyanate which are deep red—All ferric salts are reduced to ferrous by H_2S .

KOH, K_2CO_3 precipitates red brown $Fe_2(OH)_6$ soluble in acids—this pre- and $BaCO_3$: } cipitation is prevented by sugar.

H_2S :..... reduces ferric to ferrous salts, setting free S.



$(H_4N)_2S$:..... gives FeS .



$K_4FeC_6N_6$:..... gives blue $Fe_2FeC_6N_6$, "Prussian blue."

$KCNS$:..... gives red $Fe_2(CNS)_6$, organic acids hinder.

Tannic acid:.... gives black iron tannate (ink).

Ferric Acid H_2FeO_4 —

Corresponds to H_2SO_4 and H_2MnO_4 —free acid is not known as it is unstable. Potassium ferrate K_2FeO_4 is formed when iron filings are fused with KNO_3 —(see K_2MnO_4)—crystals of K salt form in dark red prisms.

Ferrous Chloride $FeCl_2$ —

The white anhydrous salt is formed when HCl gas passes over excess of heated iron—green deliquescent $FeCl_2$ is formed by the solution of Fe in HCl —oxidizes easily to Fe_2Cl_6 —all ferrous salts absorb NO .

Ferrous Bromide $FeBr_2$ —

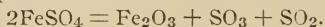
Made by action of Br on excess of Fe in the presence of H_2O —used in preparation of KBr .

Ferrous Sulphate $FeSO_4 \cdot 7H_2O$ —

"Green vitriol" is a bi-product of Cu works—obtained pure from FeS and H_2SO_4 .



Crystals effloresce in dry air, and if heated lose six molecules of water to form the colorless salt $FeSO_4 \cdot H_2O$ —at still greater heat $FeSO_4$ decomposes.



From this reaction the "fuming" or "Nordhausen" sulphuric acid is obtained.

Ammonium Ferrous Sulphate $FeSO_4 \cdot (H_4N)_2SO_4 \cdot 6H_2O$ —

A stable salt, prepared by evaporating together equal parts of $FeSO_4 \cdot 7H_2O$ and $(H_4N)_2SO_4$.

Ferrous Nitrate $Fe(NO_3)_2$ —

Unstable, green salt, made by action of HNO_3 on excess of iron, or from $Ba(NO_3)_2$ and $FeSO_4$.



Ferric Chloride Fe_2Cl_6 —

Anhydrous, by heating iron in current of Cl—hydrated $\text{Fe}_2\text{Cl}_6 \cdot 6\text{H}_2\text{O}$ by passing Cl gas into solution of FeCl_2 —anhydrous is brown—hydrated, red—both are deliquescent—soluble—volatile at high heat—forms basic salts easily—used as a reagent for organic substances.

Other $(\text{Fe})_2$ salts are formed by oxidation of the (Fe) ferrous.

CYANOGEN COMPOUNDS OF IRON.

Ferrous Cyanide $\text{Fe}(\text{CN})_2$ —

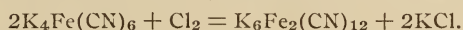
Is found as a white unstable precipitate when KCN is added to the solution of a ferrous salt—with an excess of KCN the double salt $\text{Fe}(\text{CN})_2 \cdot 4\text{KCN}$ or $\text{K}_4\text{FeC}_6\text{N}_6$ is formed.

Potassium Ferrocyanide $\text{K}_4\text{FeC}_6\text{N}_6$ —

This is known commercially as yellow prussiate of potash, and made by heating $\text{K}_2\text{CO}_3 + \text{Fe}$ with nitrogenous bodies to a red heat—a sweetish, non-poisonous salt—used in dyeing and as a test for Fe_2 and CN salts—with concentrated HCl it forms the acid, hydrogen ferrocyanide, or ferrohydrocyanic acid $\text{H}_4\text{Fe}(\text{CN})_6$, which is unstable—with ferric salts, ferrocyanide forms prussian blue.

Potassium Ferricyanide $\text{K}_6\text{Fe}_2\text{C}_{12}\text{N}_{12}$ —

Is formed when the ferrocyanide is treated with Cl.



Separates in red prisms, known as red prussiate of potash—its water solution reduces to ferrocyanide by light—its alkali solution is a strong oxidizing agent—with ferrous salts, ferricyanide forms “Turnbulls blue.”

IRON SULPHIDES.

Ferrous Sulphide FeS —

Gray, metallic mass, made by fusing Fe and S—or merely moistening a mixture of Fe and S at ordinary temperature.—Is obtained pure through precipitation from ferrous salts by $(\text{H}_4\text{N})_2\text{S}$ — FeS is soluble in acid and largely used for making H_2S .



Ferric Sulphide Fe_2S_3 —

Black sulphide analogous to Fe_2O_3 , made by heating Fe and S in proper proportions.

“Pyrite” FeS_2 —

A natural sulphide not analogous to any known oxide.

NICKEL—Ni.

At. Wt. 58.6—Val. II.

Occurrence—

Generally accompanies Co—chief ore is NiAs “Kupfernickel”—with Fe only in meteorites.

Properties—

Hard white metal—malleable—weldable—ductile—more tenacious than iron—magnetic—but slightly affected by the air, hence used to plate other metals—soluble in most acids but like Fe, is passive in HNO_3 concentrated.

Oxides and Salts—

Of nickel are closely analogous to those of Co.—Salts are green, complementary to the red of Co salts—the hydrate $\text{NiO}_2 \cdot \text{H}_2\text{O}$ is somewhat soluble in H_4NOH , with a blue color similar to Cu.

COBALT—Co.

At. Wt. 58.6. Val. II or III.

History—

The name “kobold” was given first to minerals which colored glass blue, afterwards to the metal which was discovered in them.

Occurrence—

Never in the free state, usually in combination with As and S—chief ores are “smaltite,” CoAs_2 , and “cobaltite,” $\text{CoAs}_2 \cdot \text{CoS}_2$.

Extraction—

By roasting the ore a crude arseniate is obtained, “zaffre;” this treated with KHSO_4 gives the crude oxide which is reduced with carbon to metallic Co.

Properties—

A reddish metal—lustrous—malleable—ductile—difficultly fusible—harder and more tenacious than iron—can be magnetized—not attacked by air or water—slightly by HCl and H_2SO_4 —easily by HNO_3 . Chief compounds are cobaltous, *i. e.*, where Co is bivalent.

COBALT OXIDES.

Cobaltous Oxide CoO —

Cobalt forms three oxides, CoO , Co_2O_3 , Co_3O_4 . CoO is a brown powder formed by heating Co_2O_3 with H —is the basis of cobaltous salts—with water it forms $\text{Co}(\text{OH})_2$, COBALTOUS HYDRATE, a rose colored precipitate turning brown— $\text{Co}(\text{OH})_2$ like FeO_2H_2 must be made in an atmosphere of H .

Cobaltic Oxide Co_2O_3 —

A black powder formed by gentle ignition of $\text{Co}(\text{NO}_3)_2$ —it is basis of cobaltic salts, which are very unstable.

Cobaltous-Cobaltic Oxide Co_3O_4 —

A black powder formed by strong ignition of the nitrate—it is not magnetic and forms no salts.

OTHER Co COMPOUNDS.

Cobaltous Chloride CoCl_2 —

Obtained by dissolving the oxide or carbonate in HCl —it crystallizes with $6\text{H}_2\text{O}$ to form red crystals, which, when anhydrous are blue—the red crystals are also turned blue by H_2SO_4 , which takes up its water of crystallization—characters written with the red solution are almost invisible but turn to distinct blue when heated (sympathetic ink). The Br and I compounds are similar.

Cobaltous Sulphate $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$ —

Crystallizes in dark red prisms, and resembles ferrous sulphate.

Cobaltous Nitrate $\text{Co}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$ —

Forms in red deliquescent prisms—is much used in blowpipe work, and gives characteristic blowpipe reaction with Zn (Rin-

maus green) and Al (Thenard's blue) also a blue with phosphates.

Cobalt Silicate—

No silicates are found in nature, but when glass is fused with a Co salt it forms a dark blue silicate, called "smalt"—when reduced to powder is used as a paint.

COBALT CYANIDES.

Cobalt Cyanides—

Are formed by dissolving the hydrate in KCN—in excess of KCN a double salt is formed, $\text{Co}(\text{CN})_2 \cdot 2\text{KCN}$ —when this is boiled with an oxidizing agent a cobaltcyanide, $\text{K}_6\text{Co}_2(\text{CN})_{12}$, is formed and from this KOH does not precipitate $\text{Co}_2(\text{OH})_6$.

COBALT SULPHIDES.

Cobalt Sulphides CoS —

The most common is CoS , a black precipitate, formed when $(\text{H}_4\text{N})_2\text{S}$ is added to a Co salt—insoluble in alkalies or cold dilute HCl —soluble in strong acids.

GOLD—Au.

At. Wt. 196.7 Val. I. and III.

History—

Long known to the ancients and called the king of metals—chief object of alchemists, was the transmutation of base metals to gold.

Occurrence—

Chiefly as native gold, alloyed with more or less Ag, Cu, Pb and Bi—is widely distributed, but generally occurs in small quantities—its ores are unimportant.

Properties—

Most malleable and ductile of metals—highly tenacious—at high temperature is volatile—melted has a red yellow color—on cooling, Au contracts more than other metals—no single acid dissolves Au, but with aqua regia ($\text{HCl} + \text{HNO}_3$) it forms the chloride—the other haloids also attack it, and melted KNO_3 forms the oxide.

Auric Chloride AuCl_3 —

Formed by solution of Au in Aqua regia—a reddish brown, deliquescent mass of crystals, which dissolve readily in alcohol and ether—All reducing agents are oxidized by AuCl_3 , with precipitates of gold—thus FeSO_4 precipitates gold from the chloride as a lustreless brown precipitate—and stannous chloride SnCl_2 , precipitates the oxide Au_2O as “*purple of Cassius*.”

Aurous Chloride AuCl .

When auric chloride is heated to 180° it forms aurous chloride, AuCl , a white powder in H_2O —ignited AuCl decomposes to $\text{Au} + \text{Cl}$.

Aurous Oxide Au_2O —

Formed in “purple of cassius” also when KOH acts on aurous chloride—dark violet powder—decomposed by heat—changed to Au and AuCl_3 by HCl .

Auric Oxide Au_2O_3 —

When AuCl_3 is heated with MgO a brown precipitate is formed. If excess of MgO is removed by concentrated HNO_3 auric oxide is left as a brown powder.

Auric Hydroxide—

If MgO is removed with dilute HNO_3 , $\text{Au}(\text{OH})_3$ remains as a yellow-red powder—oxide and hydroxide are insoluble in water and acids, but have acid properties and are soluble in alkalies. The hydroxide is called auric acid and forms aurates, which are derived from meta-auric acid, HAuO_2 .

Auric Sulphide Au_2S_3 —

H_2S or soluble sulphides precipitate auric sulphide, as a brown precipitate, soluble in alkaline sulphides.

Gold Cyanide AuCN —

When gold or its oxide is dissolved in KCN the colorless double salt, AuCN.KCN , may be crystalized out—from this metallic Au is easily precipitated by electrolysis, hence used in gilding.

PLATINUM—PT.

At. Wt. 194.4. Val. II. and IV.

Occurrence—

Found only in metallic state, generally alloyed with palladium, iridium and other rare metals.

Extraction—

The ore is dissolved in aqua regia, and precipitated by $(\text{H}_4\text{N})\text{Cl}$ —this forms by ignition a double salt of iridium bearing platinum in a spongy mass (*platinum sponge*.) This is used directly in making platinum vessels—if desired the iridium, which diminishes malleability but increases hardness and resistance to reagents, may be farther removed.

Properties—

Heavy, lustrous, gray-white metal—softer than Ag—not affected by air nor common acids—dissolved by aqua regia and chlorine water—attacked by fusing caustic alkalies—finely divided as in Pt sponge, and Pt black, it remarkably influences the chemical combination of gases.

PLATINUM COMPOUNDS.

Platinic Chloride PtCl_4 —

Is the most important salt, made by dissolving Pt in aqua-regia. It dissolves in water to a red yellow solution, important because it forms with H_4N or K an insoluble salt—the oxides and salts of Pt in general are formed, and act in the same way as those of Au.—Their formula is different because Pt acts with valence of two or four, thus: PtCl_2 , PtCl_4 , PtS_2 , etc—all Pt salts leave, on ignition, a residue of metallic Pt.

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